THE ELEMENTS

OF

CHEMICAL ARITHMETIC

WITH

A SHORT SYSTEM OF ELEMENTARY QUALITATIVE ANALYSIS.

BY

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PREFACE.

THIS little manual is intended to supplement the teaching of the text-books of descriptive chemistry, and to be used as a companion to them, by those who desire to make the whole subject more practical. It is the result of the author's experience after several years of elementary science teaching.

Part I. contains some of the more important rules and principles of chemical arithmetic, followed by a series of problems, which will not be found to be above the comprehension of the average student in the schools. The matter relating to chemical theory, and the rules, have been collected from the best authorities.

Part II. is devoted to an elementary system of qualitative analysis, and the best methods have been adopted. This part of the book can be used separately, and can be taught together with any good work in descriptive chemistry, such as Eliot & Storer's, Shepard's, Remsen's, or Avery's Chemistry. An intelligent student can, with the occasional supervision of his instructor, work out by himself the reactions and the separations as given in

the tables. The tables are those generally in use. Tests are given for the more common metals and acids only, and the reagents indicated are those which almost any school laboratory will afford.

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PART I.

CHEMICAL ARITHMETIC.

CHAPTER I.

INTRODUCTION.

1. Matter is anything that occupies space.

2. Divisions of Matter. Three divisions of matter are recognized in science, — masses, molecules, and atoms.

A *mass* of matter is any portion of matter appreciable by the senses.

A *molecule* is the smallest particle of matter into which a body can be divided without losing its identity; or it is the smallest portion of matter which can exist by itself.

An *atom* is a still smaller particle produced by the division of a molecule; or it is the smallest portion of matter that can go into combination.

EXAMPLES. The sun and a grain of sand are masses of matter. The smallest particle of salt which can exist and which exhibits the properties of salt is a molecule. The minute particles of chlorine and sodium which compose the molecule of salt are atoms.

A mass is made up of molecules, and a molecule is composed of atoms.

3. Attractions of Matter. The three forms of attraction admitted in science are: —

First. *Gravitation*, or the attraction between masses.

Second. *Cohesion*, or the attraction between like molecules; adhesion between unlike molecules.

Third. *Chemical attraction*, or the attraction between unlike atoms.

EXAMPLES. The attraction between the sun and the planets, or between the earth and all bodies upon it, is *gravitation*. The attraction between the molecules of a piece of marble is *cohesion*. The attraction between a liquid and solid, — as, for instance, when you dip your hand into water it becomes wet, — or between two different solids at the surface, as shown by the action of cements, is *adhesion*. The attraction between the unlike atoms of chlorine and sodium by means of which we have an entirely different substance, salt, is *chemical attraction*.

4. Province of Physics. Physics is that department of physical science which studies the results which come from the molar and molecular conditions of matter.

5. Province of Chemistry. Chemistry studies matter in its atomic condition. It investigates the laws and conditions of chemical changes, and seeks to account for some of the phenomena connected therewith.

6. Physical Changes. Physical changes are those which take place outside the molecule; they have no effect upon the molecule itself nor alter the identity of the matter operated on. The study of physics is a study of physical changes.

7. Chemical Changes. Chemical changes take place through the atoms and within the molecule. They alter the character of the molecule, and hence destroy the

identity of the matter itself. The study of chemistry is a study of chemical changes.

Examples. The change of water into ice and steam, or the change of any solid into a liquid, or of any liquid into a vapor, are *physical changes*. But when water is subjected to the influence of the electric current, it undergoes a more radical change; the water disappears, and in its place appear two gaseous substances, oxygen and hydrogen, entirely different from the water from which they were derived. This is a *chemical change*.

8. Physical Properties. Physical properties are those properties which bodies possess in virtue of their molar or molecular condition.

9. Chemical Properties. Chemical properties are those which result from the atomic composition of the molecule.

10. Chemistry defined. Chemistry is that branch of physical science which treats of the atomic composition of bodies, and of those changes in matter which result from an alteration in the kind, the number, or the relative position of the atoms which compose the molecule.

11. Analysis and Synthesis. The two processes by which the chemist seeks to find out the composition of matter are analysis and synthesis.

Analysis consists in separating the molecule into its constituent atoms.

Synthesis consists in putting together constituent atoms to form the molecule.

CHAPTER II.

MOLECULES AND ATOMS.

12. Chemical Definition of the Molecule. A molecule is the smallest particle of any substance which can exist in a free state in nature.

Molecules classified. Molecules are of two classes: —

First. *Elementary molecules*, or those whose atoms are alike.

Second. *Compound molecules*, or those whose atoms are unlike.

13. Simple Substances are those whose molecules contain like atoms.

14. Compound Substances are those whose molecules contain unlike atoms.

15. Number of Simple Substances. There are sixty-eight elementary substances, as far as has been investigated by chemical science; that is, sixty-eight substances whose molecules contain like atoms. Therefore it is obvious that there are sixty-eight different kinds of atoms. From combinations of these sixty-eight kinds of atoms all the different varieties of matter result. We cannot resolve a simple substance into any other substances or atoms.

16. Ampère's Law. "Equal volumes of all gases, simple as well as compound, under like conditions of temperature and pressure, contain the same number of molecules."

From this law, which is the most important law of modern chemistry, it results, —

First. That the molecules of all bodies in the gaseous state are of the same size.

Second. That the weight of any molecule, compared with that of hydrogen, is proportional to the weight of any given volume, also compared with the same volume of hydrogen.

EXAMPLES. If 1 liter of nitrogen, which weighs 14 times as much as a liter of hydrogen, contains the same number of molecules, then it is obvious that each molecule of nitrogen must be 14 times as heavy as a molecule of hydrogen.

17. Number of Atoms in the Molecule of Hydrogen. Assuming that 1 volume of hydrogen contains 1000 molecules, then, according to the law of Ampère, 1 volume of chlorine will contain 1000 molecules also.

Suppose these volumes (that is, 1 volume of hydrogen containing 1000 molecules and 1 volume of chlorine containing 1000 molecules) be mixed together and exposed to the action of the sunlight, they combine, forming 2 volumes of the new substance, hydrochloric acid gas, which 2 volumes, by the same law, will contain 2000 molecules. Upon analysis, each molecule of hydrochloric acid gas will be found to contain 1 atom of hydrogen and 1 atom of chlorine. That is, the 2000 molecules will contain 2000 atoms of hydrogen and 2000 atoms of chlorine. The 2000 molecules will contain, therefore, 4000 atoms; or, each molecule will contain 2 atoms. Hence each molecule of hydrogen is made up of 2 atoms.

18. Molecular Weights. If the weight of the hydrogen atom be taken as 1, then, since its molecule contains 2 atoms, its molecular weight will be 2.

Since the molecule of a compound gas or vapor occupies a volume twice as large as that occupied by the atom of hydrogen, it is obvious that the specific gravity of the gas or vapor may be found from the molecular weight by dividing the latter by 2. The *specific* gravity of a compound gas or vapor is, therefore, one-half its molecular weight. The molecular weight of any substance may be obtained by multiplying its density in the state of gas by the molecular weight of hydrogen; that is, by 2.

EXAMPLES. The density of oxygen gas, for example, is 16; that is, any volume as 1 liter weighs 16 times as much as 1 liter of hydrogen. Its molecule must be, therefore, 16 times as heavy. The molecular weight of hydrogen is 2; therefore the molecular weight of oxygen will be $16 \times 2 = 32$.

The weight of 1 liter of hydrogen is called 1 crith, and the weight of the hydrogen atom 1 microcrith.

19. Number of Atoms in the Molecule. The number of atoms in a molecule is obtained by dividing the molecular weight by the atomic weight.

EXAMPLE. The molecular weight of oxygen is 32, and its atomic weight 16. The number of atoms in the molecule is 32 divided by $16 = 2$. The molecular weight of phosphorus is 124, and its atomic weight 31; its molecule, therefore, contains 4 atoms.

Properties of Atoms.

20. Definition. An atom is the smallest particle of simple matter which can enter into the composition of a molecule.

21. Atomic Weight. The relative weight of any atom referred to hydrogen as unity is its atomic weight. It is the smallest weight of any simple substance which can take part in the formation of a chemical compound.

The molecular weight of any substance is the sum of the weights of its constituent atoms.

22. Quantivalence. The quantivalence of an atom is the quality of its combining power, expressed in hydrogen units. It expresses the number of hydrogen atoms with which it can combine or for which it can be exchanged.

Examples. The quantivalence of zinc is 2, because 1 atom replaces 2 of hydrogen. The quantivalence of carbon is 4, because 1 atom of carbon requires 4 of hydrogen to satisfy it in combination. Atoms are called monads, dyads, triads, tetrads, pentads, hexads, and heptads, according to their quantivalence. The Latin numerals are used for the adjective terms. These atoms are univalent, bivalent, trivalent, quadrivalent, quinquivalent, sexivalent, and septivalent. Atoms whose quantivalence is even are called *artiads;* those whose quantivalence is odd are called *perissads.*

An atom may form several compounds with the same substance. Therefore its quantivalence may vary. It always increases or diminishes by 2, so that it may have quantivalence of 1, 3, 5, or 7, or of 2, 4, or 6. A perissad atom can never become an artiad by such a change, nor can an artiad become a perissad.

EXAMPLES. Iron in iron sulphate is a dyad, in pyrites it is a tetrad, and in ferric acid a hexad. Chlorine forms a series of compounds with oxygen in which its quantivalence is 1, 3, 5, and 7.

Atoms are divided into two classes, according to the quality of their combining power.

First. *Positive atoms* are those which are attracted to the negative pole in electrolysis, and whose hydrates are bases.

Second. *Negative atoms* are those which go to the positive electrode, and whose hydrates are acids.

23. Atomic Symbols. Berzelius, in 1815, proposed an abbreviated form of chemical language. In this system each atom has for its symbol the first letter of its Latin name. When the names of two different atoms begin with the same letter, a second letter suggestive of the name is added.

EXAMPLES. Ag stands for an atom of silver; Fe, for an atom of iron; Sn, for one of tin, etc. (on page 89 will be found the table of the symbols of the elements).

Each atomic symbol stands not only for the atom, but represents its atomic weight.

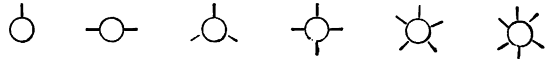
EXAMPLES. Fe (ferrum) represents 56 weight-units of iron; Hg (hydrargyrum), 200 weight-units of mercury; O, 16 weight-units of oxygen.

24. The quantivalence of an atom is indicated by placing Roman numerals above or a little to the right of the symbol. Sometimes minute-marks are used.

EXAMPLES. 1. H or H' stands for the monad hydrogen atom; 2. S or S'' stands for the bivalent sulphur atom; 3. P or P''' for the trivalent phosphorus atom; 4. C or C'''' for the quadrivalent carbon atom.

Sometimes graphic symbols are used to represent the quantivalence atoms, the graphic symbols being a circle with lines called bonds radiating from it; as, for example,

Monad. Dyad. Triad. Tetrad. Pentad. Hexad.



The circles are usually omitted, the bonds radiating from the symbol. The number of bonds and not their direction is significant, as, for example,  stands equally for 1 atom of dyad oxygen. $N\equiv$, $\overset{|}{N}=$, or $-N=$ equally represent the atom of trivalent nitrogen.

25. Multiplication of Atoms. Atoms are multiplied by placing an Arabic numeral below and to the right of the symbol.

Examples. C_2 represents 2 atoms of carbon. N_4, 4 atoms of nitrogen. Cl_3, 3 atoms of chlorine.

Molecules are multiplied by enclosing their symbols in brackets and placing the numeral outside, below, and to the right.

Examples. $(H_2)_6$ represents 6 molecules of free hydrogen. $(Br_2)_2$ stands for 2 molecules of bromine.

CHAPTER III.

COMPOUND MOLECULES AND VOLUME RELATIONS.

26. Compound Molecule. A compound molecule is one whose constituent atoms are unlike. Compound molecules are formed by the union of atoms according to the law of quantivalence.

27. Molecular Weight. The molecular weight of a compound molecule is the sum of the atomic weights of its constituents. It is always equal to twice the density of the substance in the state of gas.

28. Classification of Compound Molecules. Compound molecules are divided into two classes: first, those whose atoms are directly united, called Binaries; second, those whose atoms are indirectly united, called Ternaries. A binary compound is formed by the union of two simple substances, the termination IDE being the characteristic: as, for example, sodium and chlorine yield sodium chloride; silver and sulphur yield silver sulphide; calcium and iodine yield calcium iodide. In some cases the number of atoms of each constituent is to be indicated.

This is done by prefixing Greek numerals to each of the names given; as, for example, 1 atom of C and 2 of O form carbon dioxide, 1 atom of P and 5 of Br form phosphorus pentebromide.

29. Definition of an Acid. An acid molecule is one which consists of one or more negative atoms united by

hydrogen and oxygen. It is a compound of hydrogen and oxygen with some non-metallic element, and possesses the property of turning blue litmus paper or solution red.

30. Definition of a Base. A basic molecule is one which contains one or more positive atoms united by hydrogen and oxygen. It is a compound of hydrogen, oxygen, and some metallic element, and possesses the property of restoring the color to vegetable blues, which have been reddened by an acid.

31. Definition of a Salt. A saline molecule is one which contains a positive atom or group of atoms, united by oxygen to a negative atom or group of atoms. It is formed by the action of an acid upon a base, and since it contains no hydrogen, has no action upon vegetable colors.

32. Compound Radical. A compound radical is a group of atoms, which goes into combination like a single atom. It may be composed of two or more elements; as, for example, (NH_4) ammonium, (C_2H_5) ethyl.

33. Normal, Acid, Basic, and Double Salts. A salt is formed by substituting a metal for the hydrogen of an acid, each bond of the metal displacing one atom of hydrogen. A normal salt is formed by displacing all the hydrogen of the acid with an equivalent metal. An acid salt is formed by exchanging a part of the hydrogen of an acid for an equivalent of metal.

A basic salt is formed by the substitution of a metal in part for the hydrogen of an acid, and in part for the half or the whole of the hydrogen of water (H_2O).

Double salts are those containing two or more different positive or metal atoms.

EXAMPLES. K' with HNO_3 forms KNO_3, displacing H. KNO_3 is a normal salt, formed also by acting upon HNO_3 by KHO, as:—

$$HNO_3 + KHO = KNO_3 + H_2O.$$

K' with H_2SO_4 may form $KHSO_4$, an acid sulphate formed also by

$$KHO + H_2SO_4 = HKSO_4 + H_2O.$$

Bi''' with $\left\{ \begin{matrix} HNO_3 \\ H_2O \end{matrix} \right.$ forms $Bi \left\{ \begin{matrix} NO_3 \\ O \end{matrix} \right.$, usually written $BiONO_3$ a basic nitrate.

$NaCa''SbO_4$, sodio-calcium antimonate is an example of a double salt, or $Ba''Zn''SiO_4$, baro-zincic silicate.

Monobasic acids can form only normal salts. Polybasic acids can form normal, acid, and double salts.

34. Chemical Equations. A chemical equation is the expression in symbols of a chemical reaction, or change. The sign plus (+) indicates *added to*, and the sign minus (−), *taken from*, and the sign of equality (=), *equals to*. The equation must be a true equation; that is, the sum of the weights of the atoms on one side must equal the sum of the weights of the atoms on the other side.

35. The substances entering into the reaction are called factors; these constitute the first member. The substances issuing from the reaction are called products; these constitute the second members.

The equation, representing the reaction of two molecules upon each other may be written by the following rule:—

Place the formulas of the factors, connected by the sign plus, as the first member of the equation, and the formulas of the products, also connected by the sign plus, as the second.

EXAMPLES. Let AB and EF be two molecules. The reaction between them would be represented by the equation

$$AB + EF = AF + BE.$$

36. Weight of the Factors and Products. The quantities of matter taking part in a chemical change are definite in weight, since each formula represents a definite weight, viz., the molecular weight. For the same reason no loss of weight can be the result of any chemical reaction.

37. There are three kinds of chemical reactions: —

First. Analytical reactions; that is, the separation of a complex molecule into simpler ones.

Second. Synthetical reactions, or the union of two or more simple molecules to form a more complex one.

Third. Metathetical reactions, or the transposition or exchange of atoms between molecules.

EXAMPLES. An *analytical reaction* may be represented by the general equation

$$EF = E + F;$$

or, taking an actual example,

$$\underset{\text{Mercuric oxide.}}{HgO} = \underset{\text{Mercury.}}{Hg} + \underset{\text{Oxygen.}}{O};$$

that is, one molecule of mercuric oxide will yield one molecule of mercury and one molecule of oxygen.

Synthetical reactions may be represented by the general equation

$$E + F = EF;$$

or, taking an actual example,

$$\underset{\text{Iron.}}{Fe} + \underset{\text{Sulphur.}}{S} = \underset{\text{Iron sulphide.}}{FeS};$$

that is, one molecule of iron and one molecule of sulphur yield one molecule of iron sulphide.

Metathetical reactions may be represented by the general formula

$$AB + EF = AF + BE;$$

or, practically,

$$\underset{\text{Zinc}}{Zn} + \underset{\text{Sulphuric acid}}{H_2SO_4} = \underset{\text{Zinc sulphate}}{ZnSO_4} + \underset{\text{Hydrogen.}}{2\,H.}$$

38. The conditions which form chemical change depend upon the facility with which the atoms of any molecule may be rearranged. When substances are in the gaseous or liquid state, these changes between atoms take place most readily. Hence, fusion or solution or vaporization facilitate chemical action. Heat is therefore the great aid to the chemist.

CHAPTER IV.

STOICHIOMETRY. — CHEMICAL ARITHMETIC.

39. Definition. By Stoichiometry we mean that department of chemistry which treats of the numerical relations of atoms. The calculations of these numerical relations, whether of volume or weight, depend upon the fact that every atom has its own weight, called the atomic weight. The atomic weight is the smallest portion by weight of any simple or elementary substance — referred to the atom of hydrogen as unity — which can take part in a chemical change.

40. Unit of Weight. The weight of the hydrogen atom is called a microcrith. (The weight of one liter of hydrogen under general conditions of temperature and pressure is one crith.) We adopt the term microcrith for convenience' sake.

41. All chemical changes take place between definite quantities of matter, as represented by a chemical equation. An equation expresses not only the fact of chemical reaction between two bodies, but also indicates the quantities by weight concerned in it.

RULES.

42. From the Formula of a Substance to find its Molecular Weight. The molecular weight of a compound is the sum of the atomic weights of all the atoms of the elements which compose it.

The name of each element present being written in a column, and opposite to each the multiple of its atomic weight which is present in the compound, on adding these numbers together the molecular weight of the compound is obtained.

Thus the molecular weight of sulphuric acid, H_2SO_4, is

$$\begin{aligned} H &= 1 \times 2 = 2 \\ S &= 32 \times 1 = 32 \\ O &= 16 \times 4 = \underline{64} \\ & \qquad\qquad\;\; 98 \end{aligned}$$

43. To find the Percentage Composition of any Substance in a Molecule. *Rule.* Multiply the atomic weight by the number of atoms, and this product by 100. Divide the final product by the molecular weight, and the quotient will be the percentage amount of that constituent.

EXAMPLE. What is the percentage composition of carbon dioxide, Co_2?

$$\begin{aligned} &\text{Carbon} &= 12 \\ &\text{Oxygen, } 16 \times 2 &= \underline{32} \\ &\text{Molecular weight,} & 44 \end{aligned}$$

$$\text{Carbon} = 12 \times \tfrac{100}{44} = 27.27 \text{ per cent.}$$
$$\text{Oxygen} = 32 \times \tfrac{100}{44} = 72.73 \text{ per cent.}$$

This rule can be expressed by a general formula. Let m represent the molecular weight, a the atomic weight of any constituent, n the number of atoms, and x its percentage amount; then we have the proportion:—

$$m : an :: 100 : x,$$

from whence the formula

$$x = \frac{an \times 100}{m}. \qquad (1)$$

In the above formula, when any three of the quantities a, n, m, and x are known, the fourth can be found. Whence, to find the number of atoms of any constituent in a molecule, "x," "a," and "m" being known, we have, by transposing formula (1): —

$$n = \frac{mx}{100a}, \tag{2}$$

also

$$a = \frac{mx}{100n}, \tag{3}$$

and

$$m = \frac{an \times 100}{x}. \tag{4}$$

44. To calculate from an Equation a Mass.

Rule. Find the multiples of the atomic or molecular weights of the substances given and asked in the equation, and work the proportion.

The molecular weight of substance given : the molecular weight of substance asked : : the real mass of substance given : the real mass of substance asked. Thus, to find now many grams of sodium sulphate can be obtained from 100 grams of sodium hydrate: —

EXAMPLE 1. $2\,NaHO + H_2SO_4 = 2\,H_2O + Na_2SO_4.$

$2 \times 40 \qquad\qquad 142.$

$2 \times 40 : 142 : : 100 : x$

$x = \frac{14200}{80} = 177.5$ grams.

EXAMPLE 2. In the equation

$$KNO_3 + H_2SO_4 = HNO_3 + HKSO_4$$
$$101 \;+\; 98 \;=\; 63 \;+\; 136.$$

125 grams of KNO_3 yield 77.97 grams of HNO_3, whose molecular weight is 63. What is the molecular weight of KNO_3?

This rule can be simply expressed by the general proportion: —

$$M : m :: W : w,$$

where M represents the molecular weight of the substance given, m the molecular weight of the substance asked, W the real mass of the substance given, and w the real mass of the substance asked; whence

$$M = \frac{mW}{w}, \tag{1}$$

$$W = \frac{Mw}{m}, \tag{2}$$

$$m = \frac{Mw}{W}, \tag{3}$$

$$w = \frac{mW}{M}. \tag{4}$$

45. The Relations of Weight to Volume. 1. To find the volume occupied by a given weight of any gas. *Rule.* Divide the weight of the gas given by the weight of 1 liter; the quotient is the number of liters.

2. To find the weight of any given volume of gas. *Rule.* Multiply the number of liters of gas by the weight of 1 liter; the product is the weight of the given volume.

EXAMPLES. 1. What volume is occupied by 6.08 grams of oxygen, the weight of 1 liter of oxygen being 1.43 grams? $6.08 \div 1.43 = 4.25$ liters. *Ans.*

2. What is the weight of 25 liters of nitrogen gas, 1 liter weighing 1.26 grams? $1.26 \times 25 = 31.5$ grams. *Ans.*

46. Density of Gases. The density of any gas expresses how many times the gas is heavier than hydrogen. Knowing the density, the weight of 1 liter may readily be obtained by multiplying it by the weight of 1 liter of

hydrogen, 0.0896 grams, or 1 crith. The molecular weight of any substance being the weight of 2 volumes of the substance in the state of gas, it is evident that its density in the state of gas may be obtained by dividing its molecular weight by 2. With few exceptions, the density of any elementary gas is expressed by the same number as its atomic weight, and that of any compound gas is expressed by the same number as half its molecular weight. Thus, oxygen, $O = 16$; density, 16; or 1 liter weighs 16 criths. Ammonia, $NH_3 = 17$; density, 8.5; or 1 liter weighs 8.5 criths.

47. Relation of Gaseous Volume to Pressure. To calculate the change in volume of a mass of gas produced by a change in pressure.

Boyle's Law. The volume of a mass of gas varies inversely as the pressure upon it; or the volume of a mass of gas, multiplied by the pressure at any one time, is equal to the volume of the same mass of gas multiplied by the pressure upon it at any other time. Thus, let V equal the volume of a gas under the pressure P, and let V' equal the volume under the pressure P'; then

$$VP = V'P', \text{ or } V = \frac{V'P'}{P}.$$

If the pressure upon 1000 cc. of gas be increased from 400 mm. to 800 mm., what is the new volume?

$$V = \frac{1000 \times 400}{800} = 500 \text{ cc.}$$

48. Relation of Gaseous Volume to Temperature. *Guy Lussac's Law.* When 273 volumes of gas at 0° C. are heated, they increase by one volume for every 1° C. through which they are heated. Thus: —

273 volumes	of gas	at	0° C.	become	at	1° C.	273 + 1 volumes,
273	"	"	"	"	"	2° C.	273 + 2 "
273	"	"	"	"	"	3° C.	273 + 3 "
273	"	"	"	"	"	t° C.	273 + t "

where t expresses any number of degrees on the centigrade scale.

The coefficient of the expansion of a gas is $\frac{1}{273}$ of the volume of the gas at 0° for every degree centigrade. Hence v volumes at t° C. become at T° C.

$$v \times \frac{273 + T}{273 + t}$$

volumes; which, if V stands for the volume of the gas after change of temperature t° C. to T° C., is usually written:—

$$V = \frac{v(273 + T)}{273 + t}.$$

Example 1. Find the new volume, if 1000 cc. of gas are heated from 17° C. to 27° C. The formula is:—

$$V = \frac{1000\,(273 + 27)}{273 + 17} = \frac{1000 \times 300}{290} = 1034.8 \text{ cc.}$$

Example 2. If 1000 cc. of gas at − 23° C. are heated to 27° C., find the new volume.

$$V = \frac{1000\,(273 + 27)}{273 - 23} = \frac{1000 \times 300}{250} = 1200 \text{ cc.}$$

49. If the pressure on the gas, as well as its temperature, be changed, the above formula must be combined with the one given in (46).

$$V = \frac{v\,(273 + T)}{273 + t} \times \frac{P'}{P}.$$

EXAMPLE. If 500 cc. of gas are cooled from 39° C. to 13° C., the pressure being decreased from 800 mm. to 300 mm., find the new volume.

$$V = \frac{500\,(273 + 13)}{273 + 39} \times \frac{800}{300} = 1222.2 \text{ cc.}$$

50. Density of Gases. When the temperature of and the pressure on a gas are not mentioned, it is supposed to be at 760 mm. and 0° C. A gas under these conditions is said to be normal.

The formulæ given in this and the succeeding section only apply to normal gases; hence, when necessary, the gas under consideration must be rendered normal by using the formula: —

$$V = \frac{v \times 273}{273 + t} \times \frac{P'}{760},$$

and, conversely, the volume found by these formulæ is normal, and must be reduced to the required temperature and pressure by

$$V = \frac{v\,(273 + T)}{273} \times \frac{760}{P}.$$

In the case of gases, the liter, = 1000 cc., is taken as the unit of volume, and the mass of one liter of normal hydrogen, called a crith, = .0896 gram, is taken as the unit of mass.

The density of a gas, then, is the number of criths contained in one liter of it, measured at 0° C. and 760 mm.; or the number of times it is heavier than an equal volume of hydrogen.

Hence the mass in grams of a liter of any normal gas can be found by multiplying its density by .0896.

EXAMPLE. The density of carbon monoxide is 14; required the weight of one liter.

$$14 \times .0896 = 1.2544 \text{ grams.}$$

The density of a gas referred to air may be obtained by multiplying its density referred to hydrogen by .06926, the density of hydrogen referred to air.

EXAMPLE. Nitric oxide is 15 times as heavy as hydrogen; how many times is it heavier than air?

$$15 \times .06926 = 1.0389.$$

If the density of a gas referred to air be given, its density referred to hydrogen can be obtained by multiplying its density referred to air by 14.436.

EXAMPLE. Sulphur dioxide is 2.22 times as heavy as air; find its density and molecular weight.

$$14.436 \times 2.22 = 32.042$$

is the density referred to hydrogen, and

$$32.042 \times 2 = 64.084$$

is the molecular weight.

51. Volume and Mass of Gases. It is found by experiment that 22.32 liters of any normal gas weigh a number of grams equal to the number expressing the molecular weight of the gas. Thus: —

22.32	liters	of hydrogen	($H_2 = 2$)	weigh	2	grams.
22.32	"	" oxygen	($O_2 = 32$)	"	32	"
22.32	"	" nitrogen	($N_2 = 28$)	"	28	"
22.32	"	" chlorine	($Cl_2 = 71$)	"	71	"

This volume, 22.32 liters, is commonly spoken of as "two volumes" and expressed by the symbol □□.

Since 22.32 liters (or, if great accuracy be not required, 22.4 liters) of any gas weigh its molecular weight in grams, a liter of any gas weighs its molecular weight in grams divided by 22.32 (or 22.4); and one gram of any gas occupies 22.32 (or 22.4) liters divided by its molecular weight. Hence the mass in grams of any volume of a gas can be found by multiplying the number of liters of it by its molecular weight, and dividing by 22.32.

EXAMPLE 1. Find the mass of 250 liters of chlorine ($Cl_2 = 71$).

$$\frac{250 \times 71}{22.32} = 795.25 \text{ grams.}$$

Conversely, the volume in liters of any gas can be found by multiplying the number of grams of it by 22.32, and dividing by the molecular weight.

EXAMPLE 2. Find the volume of 225 grams of hydrogen sulphide ($H_2S = 34$).

$$\frac{225 \times 22.32}{34} = 147.7 \text{ liters.}$$

EXAMPLE 3. Find the mass of 80 liters of oxygen ($O_2 = 32$) measured at 52° C. and 740 mm. The gas must be reduced to 0° C. and 760 mm.

$$\frac{80 \times 273}{273 + 52} \times \frac{740}{760} \times \frac{32}{22.4} = \frac{8880}{95} = 93.47 \text{ grams.}$$

52. Equation and Volumes of Gases. When the volume of one gas is given, and that of another gas is asked, since each molecular weight expresses two volumes of the gas, the result may often be obtained directly.

EXAMPLE 1. What volume of hydrogen chloride is formed when 10 liters of chlorine combine with hydrogen?

$$Cl_2 + H_2 = 2\,HCl.$$
$$\square\!\square + \square\!\square = 2\ \square\!\square$$

Two volumes of chlorine form twice two volumes of hydrogen chloride; hence 10 liters of chlorine form $2 \times 10 = 20$ liters of hydrogen chloride.

EXAMPLE 2. If 10 liters of hydrogen at 15° C. are burned, what volume of steam at 300° C. is formed?

$$H_2 + O = H_2O.$$
$$\square\!\square \quad \square \quad \square\!\square$$

The volume of the steam would be equal to that of the hydrogen, if the temperatures were the same, making the correction for the change of temperature.

$$V = \frac{10\,(273 + 300)}{273 + 15} = \frac{5730}{288} = 19.895 \text{ liters.}$$

When the mass of a solid or liquid is given or asked, and the volume of a gas is asked or given, the equation can only be solved in terms of the mass of the gas.

EXAMPLE 3. How much lead sulphide can be precipitated by 17 liters of hydrogen sulphide?

$$\underset{34}{H_2S} + PbN_2O_6 = 2\,HNO_3 + \underset{239}{PbS}.$$

17 liters of H_2S weigh $\frac{17 \times 34}{22.4}$ grams.

34 grams of H_2S precipitate 239 grams of lead sulphide.

1 gram of H_2S precipitates $\frac{239}{34}$ grams of PbS.

$\frac{17 \times 34}{22.4}$ grams of H_2S precipitate

$$\frac{17 \times 34}{22.4} \times \frac{239}{34} = 181.3 \text{ grams of lead sulphide.}$$

53. Gaseous Diffusion. *Graham's Law.* "The velocity of the diffusion of any gas is inversely propórtional to the square root of its density." This law applies of course to volumes. That is, when two gases diffuse through the same apparatus for equal times under similar conditions, the volume of the one gas diffused multiplied by the square root of its density is equal to the volume of the other gas diffused multiplied by the square root of its density.

$$V\sqrt{D} = v\sqrt{d}.$$

EXAMPLE. 4 liters of hydrogen diffuse through an apparatus in 10 minutes, and 1 liter of oxygen in an equal time under similar conditions; find the density of oxygen.

$$1\sqrt{D} = 4\sqrt{1}, \qquad \therefore\ D = 16.$$

For method of determining the empirical formula of a substance from its percentage composition, and for methods of determining the relative density of solids, liquids, and gases, see Appendix, pp. 93 and 95.

EXAMPLES.

MOLECULAR WEIGHTS.

1. Find the molecular weight of (*a*) carbon monoxide, CO; (*b*) magnesia, MgO; (*c*) lime, CaO; (*d*) alumina, Al_2O_3.

2. Find the percentage of oxygen in each of the above-mentioned bodies.

3. Find the molecular weight of (*a*) nitric oxide, NO; (*b*) sodium hydrate, $NaHO$; (*c*) ferric oxide, Fe_2O_3.

4. Find the molecular weight of (*a*) zinc sulphate, $ZnSO_4 . 7\,H_2O$; (*b*) copper sulphate, $CuSO_4 . 5\,H_2O$; (*c*) sodium sulphate, $Na_2SO_4 . 10\,H_2O$.

THE VOLUME OF A MASS OF GAS.

1. 1000 cc. of gas are heated from 0° to 39°. Find the new volume.

2. 1000 cc. of gas are heated from 39° C. to 52° C. Find the new volume.

3. The pressure on 10 cc. of gas is 7 meters; if the pressure be reduced to 847 mm., what is the new volume?

4. 1000 cc. of air at 13° C. occupy what volume at 65° C.?

5. If 300 cc. of gas are measured off at 28° C., what will the volume become at −14° C.?

6. A liter of gas is heated from 14° C. to 42° C. Find the new volume.

7. A liter of air at 39° C. is cooled to −26° C. Find the new volume.

8. 50 cc. of gas at 10° C. occupy what volume at 24° C.?

9. 100 cc. of air at 12° C. are heated until they occupy 145 cc. Find the new temperature.

10. 320 cc. of gas are measured off at 91° C. and 950 mm.; what is the normal volume?

11. 542 cc. of air at 269° C. and 900 mm. are cooled to 51° C., the pressure being decreased to 666 mm. Find the new volume.

12. 546 cc. of gas at 17° C. and 760 mm. are cooled to 0° C., the pressure being decreased to 600 mm. Find the new volume.

13. A quantity of oxygen, which measures 230 liters at 14° C. and 740 mm., will measure what at 0° C. and 760 mm.?

14. 1234 cc. of normal gas are cooled to −52° C., the pressure being decreased to 617 mm. Find the new volume.

Miscellaneous Examples.

1. 50 grams of potassium chlorate are heated; what mass of oxygen is given off?

$$KClO_3 = KCl + 3\,O.$$

2. How much potassium chlorate is required to make 112 liters of oxygen?

$$KClO_3 = KCl + 3\,O.$$

3. What volume of oxygen can be obtained from 20 grams of manganese dioxide by heating it alone?

$$3\,MnO_2 = Mn_3O_4 + O_2.$$

4. What mass of oxygen can be obtained by heating 522 grams of manganese dioxide?

$$3\,MnO_2 = Mn_3O_4 + O_2.$$

5. What mass of oxygen (O_2) at 10° C. and 750 mm. would fill a globe of 15 liters capacity?

6. On heating some potassium chlorate 298 grains of potassium chloride were left. What mass of chlorate was heated, and what mass of oxygen was formed?

$$KClO_3 = KCl + 3\,O.$$

7. How much potassium chlorate is required to make 70 liters of oxygen?

8. 60 grams of mercuric oxide are heated; what volume of oxygen at 91° C. and 380 mm. is given off?

$$HgO = Hg + O.$$

9. 270 grams of mercuric oxide are heated; what volume

of oxygen at 0° C. and 760 mm. is given off? And what will the gas measure at 17° C. and 700 mm.?

$$HgO = Hg + O.$$

10. What is the mass of 13 liters of oxygen (O_2) measured at 12° C.?

11. 100 grams of steam are passed over red-hot iron. What volume of hydrogen at 10° C. and 742 mm. is formed?

$$4\,H_2O + 3\,Fe = Fe_3O_4 + 4\,H_2.$$

12. How many liters of hydrogen are obtained on dissolving 16 grams of magnesium in dilute sulphuric acid?

$$Mg + H_2SO_4 = MgSO_4 + H_2.$$

13. What volume will 1000 cc. of hydrogen at 0° C. occupy at (*a*) 15° C., (*b*) 100° C., (*c*) 300° C.?

14. What volume will 1000 cc. of hydrogen at 100° C. occupy, at (*a*) 0° C., (*b*) −20° C., (*c*) −50° C.?

15. 4 liters of hydrogen are passed over heated cupric oxide; what loss of mass does the oxide undergo?

$$H_2 + CuO = Cu + H_2O.$$

16. Find the percentage of hydrogen in (*a*) HCl, (*b*) H_2S, (*c*) NH_4.

17. How much zinc is required to obtain 100 liters of hydrogen at 91° C. and 800 mm.?

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

18. What volume of oxygen at 12° C. and 762 mm. is evolved on decomposing 10 grams of water by electricity?

$$H_2O = H_2 + O.$$

19. 500 cc. of hydrogen at 39° C. are exploded with 500 cc. of oxygen under a pressure of 332.5 mm.; what volume of which gas is left?

$$H_2 + O = H_2O.$$

20. 2 grams of steam are passed over red-hot iron; what volume of hydrogen at 10° C. and 770 mm. is formed?

$$4\,H_2O + 3\,Fe = Fe_3O_4 + 4\,H_2.$$

21. 100 grams of gypsum are heated; what volume of steam at 300° C. is given off?

$$CaSO_4,\ 2\,H_2O = CaSO_4 + 2\,H_2O.$$

22. 100 grams of each variety of sodium carbonate (*a*) Na_2CO_3, $10\,H_2O$; (*b*) Na_2CO_3, $8\,H_2O$; (*c*) Na_2CO_3, $5\,H_2O$; (*d*) Na_2CO_3, H_2O, contain how many grams of water?

23. What do 100 liters of nitrogen (N_2) weigh?

24. 146 liters of nitrogen at 17° C. and 974 mm. are heated to 51° C., the pressure being decreased to 760 mm. Find the new volume.

25. Find the percentage composition of ammonium nitrate NH_4NO_3.

26. How many pounds of nitric acid are obtained on distilling 400 pounds of sodium nitrate with sulphuric acid?

$$2\,NaNO_3 + H_2SO_4 = Na_2SO_4 + 2\,HNO_3.$$

27. How much copper is required to form 10 liters of nitric oxide?

$$3\,Cu + 8\,HNO_3 = 3\,CuN_2O_6 + 4\,H_2O + 2\,NO.$$

28. What is the mass of 270 cc. of nitrogen measured over water at 8° C. and 768 mm.?

29. What is the volume of 20 grams of ammonia (NH_3) at 12° C. and 730 mm.?

30. 25.82 liters of nitric oxide diffuse through a certain apparatus in 50 minutes; what volume of hydrogen will diffuse under same conditions?

31. The pressure on 134 cc. of air is increased from 480 mm. to 1200 mm. Find the new volume.

32. What volume of air containing 21 per cent of oxygen by volume is required to burn 248 grams of phosphorus?

$$P_2 + O_5 = P_2O_5.$$

33. How much phosphorus is required to remove the oxygen from a liter of air?

$$P_2 + O_5 = P_2O_5.$$

34. 10 liters of carbon monoxide at 14° C. and 760 mm. are required; what volume of normal carbon dioxide must be passed over red-hot carbon, and what mass of carbon is absorbed?

$$CO_2 + C = 2CO.$$

35. A substance contains carbon 20 per cent, oxygen 26.6 per cent, and sulphur 53.3 per cent. Find its formula.

36. A diamond weighing 7 grams is burnt in oxygen; what volume of carbon dioxide is formed?

$$C + O_2 = CO_2.$$

37. 10 grams of turpentine are burnt; what volume of carbon dioxide is formed?

$$C_{10}H_{16} + 28\,O = 8\,H_2O + 10\,CO_2.$$

38. How much marble must be dissolved in acid to give 20 liters of carbon dioxide at 18° C. and 740 mm.?

$$CaCO_3 + 2\,HCl = CaCl_2 + H_2O + CO_2.$$

39. How much hydrogen potassium carbonate is required to give a liter of carbon dioxide?

$$HKCO_3 + HCl = KCl + H_2O + CO_2.$$

40. What volume is occupied by 177.5 grams of chlorine (Cl_2)?

41. How much salt is required to make 28 liters of chlorine?

$$2\,NaCl + MnO_2 + 2\,H_2SO_4 = 2\,H_2O + Na_2SO_4 + MnSO_4 + Cl_2.$$

42. 500 grams of potassium chlorate are heated; how much potassium chloride is left?

43. Find the percentage composition of bleaching powder, if its formula be CaO_2Cl_2.

44. From 2078 grams of sodium chloride what volume of chlorine can be obtained?

$$2\,NaCl + MnO_2 + 2\,H_2SO_4 = 2\,H_2O + Na_2SO_4 + MnSO_4 + Cl_2.$$

45. Find the percentage composition of (*a*) galena, PbS; (*b*) zinc blende, ZnS.

46. 9.6 grams of sulphur are heated in hydrogen; what volume of hydrogen sulphide is formed?

$$S + H_2 = H_2S.$$

47. 16 liters of hydrogen diffuse through an apparatus in 100 minutes; how much sulphur dioxide (SO_2) will diffuse under the same conditions?

48. What volume of sulphur dioxide at 20° C. and 740 mm. can be obtained by the action of 20 grams of sulphuric acid upon copper?

$$Cu + 2\,H_2SO_4 = CuSO_4 + 2\,H_2O + SO_2.$$

49. 100 grams of lead form 146.45 grams of lead sulphate; find the molecular weight of sulphuric acid.

$$Pb + H_2SO_4 = H_2 + PbSO_4.$$

50. The skeleton of a man weighs 24 lbs., and contains 58 per cent of calcium phosphate ($Ca_3 2\,PO_4$). Find the quantity of phosphorus present.

51. How much phosphorus can be obtained from 169 tons of bones containing 53.7 per cent of calcium phosphate ($Ca_3 . 2\,PO_4$)?

52. Find the percentage composition of crystallized hydrogen sodium phosphate ($Na_2HPO_4 . 12\,H_2O$).

53. How many grams do 10 liters of hydrogen phosphide weigh?

54. 10 liters of hydrogen diffuse through a certain apparatus in a certain time; what volume of hydrogen phosphide (PH_3) will diffuse under similar conditions?

55. In a case of poisoning, 11.73 grains of arsenic trisulphide were found; to how much arsenic trioxide does this correspond?

$$As_2S_3 + 9\,O = As_2O_3 + 3\,SO_2.$$

56. A room 15 feet long and 10 feet wide and high is covered with a paper containing .78 grains of Scheele's green ($CuHAsO_3$) per square foot. How much arsenic is there in the room?

57. To make a kilogram of potassium hydrate, how much (*a*) potassium carbonate and (*b*) calcium hydrate is required?

$$K_2CO_3 + CaH_2O_2 = CaCO_3 + 2\,KHO.$$

58. How much dry sodium carbonate can be made from 500 kilograms of common salt?

$$2\,NaCl + H_2SO_4 = 2\,HCl + Na_2SO_4.$$
$$Na_2SO_4 + CaCO_3 + 4\,C = 4\,CO + CaS + Na_2CO_3.$$

PART II.

ELEMENTARY QUALITATIVE ANALYSIS.

SECTION I.

PRELIMINARY EXAMINATION OF SINGLE SALTS CONTAINING ONE ACID AND ONE BASE.

TABLE I. **Examination for Acid.***

54. If the substance is a solid, powder a portion of it, and heat in a glass tube sealed at one end. Notice whether water is given off. If so, test its action with litmus paper.

Acid reaction indicates *sulphites*, *chlorides*, etc.

Alkaline reaction indicates salts of *ammonium* (NH_4).

If a sublimate forms: yellow (or in red globules) indicates *sulphur;* white, *ammonium salts*, *mercury*, *arsenic*, *antimony*.

Metalic mirror indicates *arsenic.*

If a gas is evolved: oxygen indicates *chlorates*, *nitrates*, *peroxides;* carbon monoxide indicates *oxalates;* nitrogen tetroxide indicates *nitrates;* ammonia indicates *ammonium salts;* carbon dioxide indicates *carbonates.*

If the substance alters in color: black indicates *organic matter;* yellow while hot indicates *zinc oxide.*

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* **Mercury, sulphur, ammonia, though not acids, are included.**

Take another portion of substance under analysis, and add HCl. Notice whether a gas is evolved with effervescence.

If it smell like burning sulphur, it indicates *sulphites*, or *hyposulphites*.

If it has the odor of rotten eggs, *sulphides*.

If it has the odor of bitter almonds, *cyanides*.

If it has the odor of chlorine on heating, *peroxides*, *chromates*, or *hypochlorites*.

If it renders lime-water turbid, *carbonates*.

Take another portion of substance, and try if it is soluble in water; if so, add $BaCl_2$ solution to a portion of the solution and notice whether a precipitate form.

A white precipitate insoluble in HCl indicates *sulphates*.

White and soluble in HCl indicates *phosphates*, *silicates*, *oxalates*, *borates*, and *fluorides*, also *carbonates* and *sulphites*.

Yellow indicates *chromates*.

If $BaCl_2$ gives no precipitate, add $AgNO_3$ to another portion of the solution and notice if a precipitate form.

White precipitate indicates *chlorides*, also *cyanides*.

Yellowish-white indicates *bromides* and *iodides*.

Black indicates *sulphides*.

In case neither water nor HCl has dissolved the substance, try HNO_3.

If this does not dissolve it, try aqua regia; and if that fails, try method described in Table II.

If the substance is dissolved in HNO_3 or aqua regia, it must be evaporated to dryness with HCl before proceeding to the examination for base.

SECTION II.

Examination for Base.

55. Having obtained a solution, add HCl. If it produces a precipitate, it indicates *silver*, *lead*, or *mercurous salts*. Add $HCl + H_2S$. If it produces a precipitate,

Black indicates *mercuric salts*, *lead*, *bismuth*, or *copper;*
Yellow indicates *arsenic*, *stannic salts*, or *cadmium;*
Orange indicates *antimony;*
Brown indicates *stannous salts*.

If $(NH_4)HO + (NH_4)Cl + (NH_4)_2S$ produce a precipitate, it indicates

Black, *iron*, *nickel*, *cobalt;*
White, *zinc* or *aluminum;*
Flesh-colored, *manganese;*
Green, *chromium*.

If $(NH_4)HO + (NH_4)Cl + (NH_4)_2CO_3$ produce a precipitate, it indicates

Barium (tinges flame green),
Strontium (tinges flame crimson),
Calcium (tinges flame dull red).

If the solution is not precipitated by any of the above reagents, it indicates *magnesium*, *potassium*, *sodium*, *ammonium*, of which the following are the individual tests: —

Magnesium is precipitated by $Na_2HPO_4 + (NH_4)HO$, white.

Potassium is precipitated (except in very dilute solu-

tions) by $PtCl_4$, precipitate insoluble in alcohol; also tinges the flame violet.

Sodium is precipitated by H_2SiF_6; also tinges flame intense yellow, not visible through blue glass.

Ammonium-salts heated with NaHO give smell of NH_3.

Table II.

Examination of Insoluble Substances. The following substances are, under certain circumstances, insoluble in acids, and must be examined specially:—

Silica, *Silicates*.
Alumina, *Aluminates*.
Oxides of *Antimony*, *Chromium*, and *Tin*.
Chrome Iron Ore.
Sulphates of *Barium*, *Strontium*, and *Lead*.
Certain *Fluorides* (*e.g.* of *Calcium*).
Certain *Sulphides* (*e.g.* of *Lead*).
Chloride, *Bromide*, and *Iodide of Silver*.
Carbon.
Sulphur.

Heat the substance in a dry tube as before, and notice if it fuses and volatilizes completely. If it smells of SO_2, it indicates *sulphur*.

If it fuses, but does not volatilize, indicates *chloride*, *bromide*, or *iodide of silver* (also will yield metallic silver on fusing on charcoal with Na_2CO_3).

If it is infusible, but disappears on heating, *carbon* (deflagrates when heated with KNO_3).

If it is infusible, but darkened in color while hot, regaining its color on cooling, *tin dioxide and antimony pentoxide*

(confirmed by blow-pipe test — tin bead malleable; antimony bead, brittle).

Notice whether. It yields a green bead with borax or microcosmic salt; it indicates *chromium oxide*, or *chrome iron ore*. It swims undissolved in a bead of microcosmic salt, *silica* and *silicates* (fuse with four times its weight of a mixture of K_2CO_3 and Na_2CO_3. Allow to cool, dissolve in water, add HCl, and evaporate to dryness. Silica will separate out as a gelatinous mass).

It yields a colorless bead, with microcosmic salt, *alumina*. (Heated on charcoal, and moistened with $CO(NO_3)_2$ and reheated, it yields a blue, infusible mass.)

It is white and infusible, but quite unaltered by heating.

Lead sulphate yields, when heated with Na_2CO_3 in blow-pipe reducing flame, malleable metallic bead.

Barium sulphate fused with Na_2CO_3 yields $BaCO_3$. Boil the fused mass with water, filter and wash; the residue dissolved in HCl yields $BaCl_2$ (flame color green), precipitated by $SrSO_4$ solution.

Strontium sulphate fused with Na_2CO_3 yields $SrCO_3$. Boil the fused mass with water, filter and wash; the residue dissolved in HCl yields $SrCl_2$ (flame color crimson), precipitated by $CaSO_4$ solution.

Calcium fluoride heated with H_2SO_4 yields HF, which etches glass.

It is black and infusible, and yields a malleable metallic bead when fused with Na_2CO_3 in the blow-pipe flame.

Lead sulphide (bead leaves mark on paper), and when dissolved in HNO_3 gives a white precipitate on addition of H_2SO_4.

The action of strong H_2SO_4 often affords a valuable indi-

cation of the nature of a salt, whether soluble or insoluble.

Thus evolution of

Sulphur dioxide indicates *sulphites* or *hyposulphites.*

Sulphuretted hydrogen indicates *sulphides.*

Hydrocyanic acid indicates *cyanides.*

Oxygen indicates *peroxides*, *chromates*, *permanganates.*

Carbon dioxide indicates *carbonates.*

Carbon monoxide indicates *oxalates*, *formates*, *ferrocyanides.*

Chlorine indicates *hypochlorites.*

Hydrochloric acid indicates *chlorides.*

Hydrofluoric acid indicates *fluorides.*

Nitric acid indicates *nitrates.*

Acetic acid indicates *acetates.*

Chlorine tetroxide indicates *chlorates.*

REACTIONS OF THE COMMONLY OCCURRING METALS WITH THE METHODS OF SEPARATION.

Grouping of the Metals.

56. The metals are divided into five groups, according to their behavior with certain substances termed *group reagents.*

Group I. (*Silver Group.*)

Group reagent, HCl. Metals whose chlorides are insoluble in water. They are precipitated from the solutions of their salts by the first group reagent, *hydrochloric acid.*

Silver, *mercury* (mercurous salts), *lead.*

Group II. (*Copper Group.*)

Group reagent H_2S in presence of HCl. Metals which in acid solutions form insoluble sulphides, are precipitated from their acidulated solutions by the second group reagent H_2S (hydrosulphuric acid).

Arsenic, *antimony*, *tin*, *lead*, *bismuth*, *copper*, *cadmium*, *mercury* (*mercuric salts*).

The three metals, *arsenic*, *antimony*, and *tin*, form a sub-group, as their sulphides are soluble in $(NH_4)_2S_2$, whilst the sulphides of the remaining metals are insoluble in that reagent.

Group III. (*Iron Group.*)

Group reagent $(NH_4)_2S$ in presence of $(NH_4)Cl$ and $(NH_4)HO$.

Metals whose sulphides and hydroxides are insoluble in water, but decomposed by dilute acids, are precipitated from neutral solutions by the third group reagent, ammonium sulphide. *Aluminium*, and *chromium* are precipitated as hydrates; the others as sulphides. *Iron*, *nickel*, *cobalt*, *zinc*, *manganese*, as *sulphides*.

Group IV. (*Barium Group.*)

Group reagent $(NH_4)_2CO_3$ in presence of $(NH_4)HO$ and $(NH_4)Cl$.

Metals whose carbonates are insoluble in water, and are precipitated from their solutions by the fourth group reagent, ammonium carbonate; *barium*, *strontium*, *calcium*.

Group V. (*Potassium Group.*)

Metals not precipitated by any of the above group reagents, as their chlorides, sulphides and carbonates are

soluble in water. They are, therefore, distinguished by individual tests: *magnesium*, *potassium*, *sodium*, *ammonium*.

57. Each group reagent will precipitate the metals of preceding groups. The metals distinguished by being insoluble as chlorides (Group I.) are also insoluble as sulphides (with Groups II. and III.) and as carbonates (with Group IV.). The second group sulphides are precipitated both from acid and from neutral solutions, though the third group sulphides are precipitated from neutral, but not from acid solutions, and second and third group metals form insoluble carbonates, as well as those of Group IV.

In the work of analysis, the first group metals may be worked with the second, but thereafter the metals found in each group must be completely removed before testing for the next group.

After filtering out a group precipitate, the reagent which produced it should be again carefully applied, with the proper conditions, to the filtrate before testing it for the next group.

The student should at first have several metallic salts given to him, and be asked merely to determine to which of the above groups each salt belongs.

He ought next to make himself familiar with the individual tests for each metal which follows, and then proceed to the separations of the different metals. It will also be well for him to attempt to frame a table of separations for each group before consulting those given in the book.

58. Reactions of the Metals of the Silver Group (Group I.).

Silver, Ag′, 108. *Solution for Reactions*, $AgNO_3$.

1. HCl produces a white, curdy precipitate of AgCl, insoluble in hot water, soluble in NH_4HO and in KCN; reprecipitated by HNO_3; darkens on exposure to light.

2. H_2S or $(NH_4)_2S$ produces a black precipitate of Ag_2S, soluble in boiling HNO_3, with separation of sulphur.

3. NaHO produces a light-brown precipitate of Ag_2O, soluble in $(NH_4)HO$.

4. $(NH_4)HO$ produces (from neutral solutions only) brown Ag_2O, soluble in excess of reagent.

5. K_2CrO_4 produces a dark-red precipitate of Ag_2CrO_4, soluble in hot HNO_3 or in $(NH_4)HO$; this solution deposits on cooling an acid chromate in needle-shaped crystals.

6. KI precipitates yellowish AgI.

7. Cu and some other metals precipitate metallic Ag.

8. Na_2HPO_4 precipitates yellow Ag_3PO_4.

9. Heated on charcoal with Na_2CO_3, in the reducing flame of the blow-pipe, yields bright, malleable metallic beads, soluble in HNO_3.

Characteristic reaction, 1.

Mercury, Hg″, 200 (*Mercurous Salts*). *Solution for Reactions*, $Hg_2N_2O_6$.

1. HCl produces white precipitate of Hg_2Cl_2 (calomel), insoluble in cold HNO_3; blackened by $(NH_4)HO$, from formation of $Hg_2Cl(NH_2)$.

2. H_2S or $(NH_4)_2S$ produces a black precipitate of Hg_2S, not dissolved by boiling HNO_3.

3. NaHO precipitates black Hg_2O, insoluble in excess of NaHO or $(NH_4)HO$; decomposes readily into HgO and Hg.

4. $SnCl_2$ precipitates gray Hg. If the fluid be poured off and the residue boiled with HCl, distinct globules are obtained.

5. KI precipitates dark-green Hg_2I_2.

6. A drop of a metal or only slightly acid solution of a mercurous salt placed on a bright copper coin will deposit mercury, and the stain will become bright by rubbing.

7. Heated in small tube with $NaHCO_3$, yields gray deposit of Hg. Hg is volatile, and condenses on the cooler parts of the tube; soluble in HNO_3.

Characteristic reactions, 1, 7.

Lead, Pb'', 207. *Solution for Reaction*, PbN_2O_6.

1. HCl precipitates (incompletely) white $PbCl_2$, soluble in boiling H_2O, or in large quantity of cold H_2O; converted into a basic salt on adding $(NH_4)HO$, without change of appearance. If $PbCl_2$ be dissolved in boiling H_2O, it will crystallize from this solution on cooling.

2. H_2SO_4 precipitates heavy white $PbSO_4$, soluble in NaHO or ammonium tartarate. This precipitate in dilute solutions only appears on standing. If there is no immediate precipitation, concentrate the solution by evaporation. $PbSO_4$ is soluble in boiling HCl, and the solution, on cooling, deposits needle-shaped crystals of $PbCl_2$.

3. H_2S or $(NH_4)_2S$ precipitates black PbS, soluble in hot HNO_3.

4. K_2CrO_4 precipitates bright-yellow (chrome-yellow) $PbCrO_4$; soluble in NaHO; soluble with difficulty in HNO_3.

5. KI precipitates bright-yellow PbI_2, soluble in boiling H_2O; the solution on cooling deposits the salt in brilliant, golden hexagonal crystals.

6. Zn precipitates metallic Pb in crystalline form. Known as the "lead tree."

7. Heated on charcoal with $NaHCO_3$, yields malleable beads of Pb, soluble in HNO_3, and at the same time a yellow incrustation of PbO on the charcoal.

Characteristic reactions, 2, 4, 5.

59. Table II. for the Separation of Silver (Group I.).

Silver, Mercury, and Lead.

(*a*) Add HCl until no further precipitation takes place. Filter from the precipitated chlorides.

Precipitate contains	Filtrate contains
AgCl, Hg_2Cl_2, $PbCl_2$.	Group II., III., IV., and V.

(*b*) Wash precipitate twice with cold H_2O, and add washings to filtrate; then twice with hot H_2O. Test part of this for Pb with dilute H_2SO_4. White precipitate indicates *lead.* Boil remainder down to obtain crystals of $PbCl_2$. If Pb is found, the precipitate is washed repeatedly with hot H_2O, till free from it. Residue indicates AgCl, Hg_2Cl_2, insoluble in hot H_2O.

(*c*) To residue add warm $(NH_4)HO$; filter.

Residue is Hg.	Filtrate Ag.
If residue is black, it indicates *mercury.* Dissolve in $HCl + HNO_3$, and test with $SnCl_2$.	Add HNO_3; white precipitate indicates the presence of *silver.*

Reactions of the Metals of the Copper Group.

60. Metals whose sulphides are insoluble in HCl and are precipitated in presence of that acid by the group reagent H_2S.

MERCURY, LEAD, BISMUTH, COPPER, CADMIUM, ARSENIC, ANTIMONY, AND TIN.

Sub-Group A. Sulphides of the above metals insoluble in $(NH_4)_2S_2$, viz.: *Mercury*, *Lead*, *Copper*, *Bismuth*, and *Cadmium*.

Mercury, Hg'' (*Mercuric Salts*). *Solution for Reactions*, $HgCl_2$.

1. H_2S produces, when added by degrees, first a white precipitate, which changes to orange, then to brownish-red, and finally to a black precipitate of HgS. These successive changes of color on the addition of H_2S are exceedingly characteristic. This precipitate is insoluble in HCl and in HNO_3, even on boiling; it is soluble in KHS and in aqua regia.

2. KHO produces a yellow precipitate of HgO insoluble in excess, except when added to very acid solutions.

3. $(NH_4)HO$ produces a white precipitate of $HgCl(NH_2)$ (white precipitate).

4. $SnCl_2$ when added in small quantities precipitates white Hg_2Cl_2; when added in excess, gray metallic Hg, which may be united into a globule by boiling with HCl.

5. KI precipitates bright-red HgI_2, soluble in excess of either KI or $HgCl_2$.

6. K_2CrO_4 precipitates an orange basic chromate easily soluble in HNO_3.

7. Reactions 6 and 7 for mercurous salts are also produced with mercuric.

Characteristic reactions, 1, 4.

Lead, Pb'', 207. *Solution for Reactions*, PbN_2O_6.

1. H_2S precipitates black PbS, even in solutions of $PbCl_2$, so that lead belongs to both the silver and the copper groups.

2. Reactions 2, 3, 4, 5, for lead in Group I. are also applicable in this group.

Bismuth, Bi''', 210. *Solution for Reactions*, $BiCl_3$.

1. H_2S precipitates black Bi_2S_3, insoluble in KHS and KHO, but soluble in HNO_3.

2. KHO or $(NH_4)HO$ precipitates white BiO.OH, which on boiling becomes yellow (Bi_2O_3); precipitate is insoluble in excess of either reagent.

3. H_2O, when added in considerable quantity to normal salts of bismuth, precipitates white basic salt of bismuth, BiOCl, insoluble in tartaric acid. Solutions of bismuth salts containing much free acid do not give this reaction with H_2O until the excess of acid has been expelled by evaporation.

4. Zn or Fe precipitates spongy Bi.

5. K_2CrO_4 precipitates yellow $Bi_2.3(CrO_4)$, soluble in HNO_3, and insoluble in NaHO.

6. Heated on charcoal with $NaHCO_3$, in reducing flame of blow-pipe, compounds of Bi yield brittle metallic globules; also a yellow incrustation of Bi_2O_3 on the charcoal. Bi is soluble in HNO_3 or aqua regia.

Characteristic reactions, 3, 6.

Copper, Cu'', 63.5. *Solution for Reactions*, $CuSO_4$.

1. H_2S precipitates black, CuS, soluble in HNO_3; insoluble in KHS, and only slightly soluble in $(NH_4)_2S_2$. CuS is also soluble in KCN, but insoluble in hot dilute H_2SO_4.

2. KHO precipitates a pale-blue $Cu(HO)_2$, insoluble in excess. If KHO be added in excess and the mixture boiled, the precipitate becomes black.

3. $(NH_4)HO$ precipitates, when added in small quantities, greenish-blue basic salt, soluble in excess of $(NH_4)HO$, forming a dark-blue solution which consists of double basic salt of copper and ammonium.

4. $K_4Fe(CN)_6$ precipitates brown $Cu_2Fe(CN)_6$, insoluble in dilute acids, but decomposed by KHO.

5. Fe precipitates Cu in the metallic state, especially in the presence of a little free acid.

6. Zn also precipitates copper solutions.

7. K_2CrO_4 precipitates a brownish-red basic chromate, soluble in HNO_3 and in $(NH_4)HO$.

8. Compounds of Cu, when heated in Bunsen flame, impart a green color, especially after addition of AgCl.

9. Heated on charcoal with $NaHCO_3$ in reducing flame, yields brittle metallic globules of bright-red color, soluble in HNO_3 or concentrated H_2SO_4.

Characteristic reactions, 3, 5, 6.

Cadmium, Cd'', 112. *Solution for Reactions*, CdN_2O_6.

1. H_2S precipitates yellow CdS, soluble in HNO_3, insoluble in KHS, KCN, and $(NH_4)_2S$. CdS is dissolved by hot dilute H_2SO_4.

2. KHO precipitates $Cd(HO)_2$, insoluble in excess of reagent.

3. $(NH_4)HO$ precipitates $Cd(HO)_2$; soluble in excess of reagent.

4. Zn precipitates Cd in brilliant scales.

5. Heated on charcoal with $NaHCO_3$ in the reducing flame, yields a brown incrustation of CdO. Cd dissolves readily in HNO_3.

Characteristic reactions, 1, 5.

TABLE III.

61. Copper Group (II.).

Separation of Mercury, Lead, Bismuth, Copper, and Cadmium (Sub-Group A).

To the filtrate from the Silver Group add an equal bulk of HCl, boil down nearly to dryness, dilute with H_2O, and pass H_2S through the hot solution. Filter.

RESIDUE.	FILTRATE.
HgS, PbS, Bi_2S_3, CuS, CdS, SnS, SnS_2, Sb_2S_3, and As_2S_3.	Groups III., IV., and V.

Wash with hot H_2O containing H_2S, until free from HCl; digest residue* with $(NH_4)_2S_2$ for about fifteen minutes. Filter.

RESIDUE.	FILTRATE.
HgS, PbS, Bi_2S_3, CuS, and CdS.	Sub-Group B. (Sn, Sb, and As.) (See Table IV.)

* In the absence of the sub-group omit this, and treat with HNO_3 as below.

Wash with hot H_2O till no longer alkaline; add a small quantity of boiling HNO_3, pouring it on several times. Filter.

Residue.

HgS.

Dissolve in aqua regia, boil down to expel acid, and test with $SnCl_2$. White precipitate changing to gray indicates **Mercury**. Confirm by reduction test, or Bunsen's test for mercury.

Filtrate.

Pb, Bi, Cu, Cd.

Add H_2SO_4, and boil down. White precipitate indicates **Lead**. Filter from $PbSO_4$, and add $(NH_4)HO$ to filtrate. Filter.

Residue.

Bi.

A white precipitate. Dissolve in HCl, evaporate to a small bulk, and add to H_2O. A white precipitate indicates **Bismuth**.

Filtrate.

Cu, Cd.

A blue solution indicates **Copper**. Adopt Method I. or II.

Method I.

Add to the blue solution KCN till colorless; pass a few bubbles of H_2S through the solution. Yellow precipitate indicates **Cadmium**. (Excess of H_2S must be avoided, because traces of mercury may be present.)

Method II.

Add HCl till acid; pass H_2S through; filter; wash rapidly and thoroughly, and treat with hot dilute H_2SO_4. Black residue indicates **Copper**. To filtrate, add H_2S. Yellow precipitate indicates **Cadmium**.

62. Sub-Group B. Sulphides soluble in $(NH_4)_2S_2$, viz.: those of *Tin*, *Antimony*, and *Arsenic*.

Tin, Sn'', 118 (*Stannous Salts*). *Solution for Reactions*, $SnCl_2$.

1. H_2S precipitates dark-brown SnS, soluble in yellow $(NH_4)_2S_2$ (nearly insoluble in colorless $(NH_4)_2S$); re-precipitated as yellow SnS_2 by HCl.
2. KHO precipitates $Sn(HO)_2$, soluble in excess of reagent.
3. $(NH_4)HO$ precipitates $Sn(HO)_2$, not soluble in excess.
4. $HgCl_2$ precipitates at first white Hg_2Cl_2; on boiling with excess of reagent, gray Hg.
5. $AuCl_3$, on addition of a little HNO_3, precipitates purple (royal purple of Cassius).
6. Zn precipitates metallic tin.
7. Mixed with $NaHCO_3+KCN$, and heated on charcoal in reducing flame, yields small globules of Sn and a white incrustation of SnO_2.

Stannic Salts, Sn^{iv}. *Solution for Reactions*, $SnCl_4$.

Stannous salts are converted into stannic by oxidizing agents, HNO_3, Cl, Fe_2Cl_6, etc.

1. H_2S precipitates yellow SnS_2, soluble in $(NH_4)_2S$, in KHO, and in boiling concentrated HCl. It is with difficulty soluble in $(NH_4)HO$, and insoluble in $(NH_4)_2CO_3$.
2. KHO or $(NH_4)HO$ precipitates white $SNO(HO)_2$, soluble in an excess of precipitant.
3. Zn produces same reaction as with stannous salts.
4. The blow-pipe reaction for stannic is the same as for stannous.

Characteristic reaction: stannous, **1, 4**; stannic, **1.**

63. *Antimony*, Sb''', 122.3. *Solution for Reactions*, $SbCl_3$.

1. H_2S precipitates orange Sb_2S_3, soluble in $(NH_4)_2S$, NaHO, and in hot concentrated HCl; insoluble in $(NH_4)_2CO_3$.

2. KHO precipitates Sb_2O_3, soluble in excess of reagent.

3. $(NH_4)HO$ precipitates Sb_2O_3, insoluble in excess.

4. H_2O produces in solutions of $SbCl_3$ a white precipitate of SbOCl, soluble in tartaric acid.

5. Zn in presence of HCl and platinum precipitates Sb as a black powder, which adheres to the platinum. The black stain on the platinum is not removed by HCl, but is immediately dissolved by warm HNO_3.

6. (Marsh's Test.) If a solution of Sb be placed in a flask in which hydrogen is being generated, SbH_3 is given off as a gas, which is decomposed by heat, Sb being deposited. This is best done by holding in the SbH_3 flame a piece of cold porcelain; a dull-black stain of metallic Sb will be deposited on it. Add to the stain on the porcelain a drop of NaClO; the stain will remain undissolved.

7. Heated with $NaHCO_3$ on charcoal in the reducing blow-pipe flame, yields brittle globules of the metal and a white incrustation of Sb_2O_3 on the charcoal.

Characteristic reactions, 5, 6.

64. *Arsenic*, As''' (*Arsenious Compounds*). *Solution for Reactions*, Na_3AsO_3.

1. H_2S (in acid solutions) precipitates yellow As_2S_3, soluble in alkaline sulphides in KHO, in HNO_3, and in $(NH_4)_2CO_3$, but nearly insoluble in boiling concentrated HCl.

2. $(NH_4)HO$ and NaHO produce no precipitates.

3. $AgNO_3$ produces in neutral solutions a pale-yellow precipitate of Ag_3AsO_3.

4. $CuSO_4$ added to a neutral solution precipitates $CuHAsO_3$ (Scheel's green), soluble in $(NH_4)HO$.

5. Cu added to an HCl solution of arsenic becomes coated with a gray film of metallic As. *Reinsch's test.*

6. Proceed exactly as in Marsh's test for Sb, substituting a solution of As for one of Sb, and observe the bluish flame with which the mixture of H and AsH_3 burns, and also the production of white fumes, As_2O_3. Obtain stains on cold porcelain as in the case of Sb. Observe the distinction in color of the stains. Dark-brown or almost black, in the case of Sb; pale-brown and lustrous in the case of As. Add to one of the stains on porcelain a drop of NaClO; it will be rapidly dissolved.

7. Arsenious compounds are converted into arsenic compounds by oxidizing agents.

Arsenic Compounds, As^v. *Solution for Reactions*, Na_3AsO_4.

1. H_2S in warm solutions free from HNO_3 precipitates As_2S_3 and S.

2. See 2, under Arsenious Compounds.

3. $AgNO_3$ precipitates from neutral solutions reddish-brown Ag_3AsO_4.

4. $MgSO_4$ in presence of $(NH_4)HO$ and $(NH_4)Cl$ precipitates white crystalline $MgNH_4AsO_4$.

5. Heated on charcoal with Na_2CO_3, all compounds of arsenic are reduced to As, which volatilizes with characteristic garlic odor.

6. See 6, under Arsenious Compounds, Marsh's test.

Characteristic tests: arsenious, 4, 5, 6; arsenic, 5, 6.

TABLE IV.

Group II.

65. *Separation of Arsenic, Antimony, and Tin (Sub-Group B).*

Solution in $(NH_4)_2S_2$ contains sulphides of As, Sb, and Sn. Add HCl until acid: the metals are re-precipitated as sulphides. Filter; wash precipitate with hot water till free from HCl; digest precipitate with one or two pieces of solid $(NH_4)_2CO_3$ and H_2O. Filter.

RESIDUE.	FILTRATE.
SnS_2, Sb_2S_3.	As.
Wash and dissolve in strong boiling HCl, dilute with water, and add to the solution a piece of zinc and platinum: Sb forms a black stain on the platinum. Dissolve in HNO_3, boil down to a small bulk, and pass H_2S through the solution: orange precipitate indicates presence of **Antimony**. The Sn deposits on the zinc; dissolve in HCl, boil down, and test with $HgCl_2$. White precipitate indicates **Tin**.	Add HCl until acid; wash precipitated sulphide, and dissolve in HCl and a little $KClO_3$, boil down to a small bulk, and apply Marsh's test. Metallic mirror, yielding octahedral crystals on heating, indicates **Arsenic**. Dissolve in H_2O, and confirm by adding $AgNO_3$ and dilute $(NH_4)HO$, to obtain yellow precipitate of Ag_3AsO_3.

66. Reactions of the Metals of the Iron Group (Group III.).

Metals whose sulphides and hydrated oxides are insoluble in water, and are precipitated on addition of the group reagent $(NH_4)_2S$ in presence of $(NH_4)HO$ and $(NH_4)Cl$.

IRON, NICKEL, COBALT, ZINC, ALUMINUM, MANGANESE, AND CHROMIUM.

Iron, Fe'', 56 (*Ferrous Salts*). *Solution for Reactions*, $FeCl_2$.

1. $(NH_4)_2S$ precipitates black FeS, soluble in HCl, insoluble in alkalies.

2. $(NH_4)HO$ or KHO precipitates white $Fe(HO)_2$, which rapidly acquires a dirty-green color, and ultimately a reddish-brown color, owing to absorption of oxygen and conversion into ferric hydrate $Fe_2(HO)_6$.

3. $(NH_4)_2CO_3$ or Na_2CO_3 precipitates white $FeCO_3$, which rapidly darkens in color.

4. $K_4Fe(CN)_6$ precipitates white $K_2Fe_2(CN)_6$, which rapidly becomes blue by oxidization to $Fe_5(CN)_{12}$ (Prussian blue).

5. $K_3Fe(CN)_6$ precipitates "Turnbull's blue," $Fe_3Fe_2(CN)_{12}$.

6. KCNS produces no coloration.

7. $BaCO_3$ produces no precipitate in cold solution of ferrous salts.

8. Ferrous compounds are converted into ferric by oxidizing agents, such as HNO_3, $KClO_3$, HCl, Cl.

9. Fused with borax in the oxidizing flame, yellowish-red beads are produced; in the reducing flame the beads become green.

Ferric Salts, Fe^{iv}. *Solution*, Fe_2Cl_6.

1. H_2S in acid solutions produces a precipitate of sulphur, and the salt is reduced to protosalt:—

$$Fe_2Cl_6 + H_2S = 2\,FeCl_2 + 2\,HCl + S.$$

2. $(NH_4)_2S$ precipitates black FeS and S, soluble in HCl and HNO_3.

3. $(NH_4)HO$ or KHO precipitates reddish-brown $Fe_2(HO)_6$, insoluble in excess of reagent.

4. $K_4Fe(CN)_6$ precipitates "Prussian blue," $Fe_5(CN)_{12}$, insoluble in HCl, soluble in $C_2H_2O_4$.

5. $K_3Fe(CN)_6$ produces a reddish-brown color.

6. KCNS produces, even in dilute solutions, an intense blood-red color, forming a soluble iron sulphocyanide. HCl does not destroy coloration, but it is destroyed by $C_2H_3O_2Na$, $HgCl_2$, H_3PO_4, and $C_4H_6O_6$.

7. $BaCO_3$ precipitates ferric solutions completely as $Fe_2(HO)_6$ mixed with basic salt.

8. The blow-pipe reactions are the same as for ferrous compounds.

9. Heated on charcoal with Na_2CO_3, compounds of Fe yield magnetic particles, but no globules.

Characteristic reactions: ferrous salts, 5; ferric, 4, 6.

67. *Nickel*, $Ni^{\prime\prime, iv}$, 58.7. *Solution*, NiN_2O_6.

1. $(NH_4)_2S$ precipitates black NiS, slightly soluble in excess, forming a brown solution, from which NiS is precipitated on boiling, insoluble in HCl, soluble in HNO_3 and aqua regia.

2. (Na)HO or KHO precipitates light-green $Ni(HO)_2$, insoluble in excess of the reagent.

3. $(NH_4)HO$ produces a precipitate of $Ni(HO)_2$, readily soluble in excess, yielding a blue fluid. Acid solutions, or those containing salts of ammonia, yield no precipitate with $(NH_4)HO$.

4. KCN precipitates yellowish-green $Ni(CN)_2$, soluble in excess and reprecipitated by HCl or H_2SO_4, and if boiled with a strong solution of NaClO, yields a black precipitate of $Ni_2(HO)_6$.

5. KNO_3 in presence of $C_2H_4O_2$ produces no precipitate.

6. Fused with borax compounds of Ni, yields reddish-yellow beads when hot in oxidizing flame. In reducing flame the bead becomes gray.

Characteristic reactions, 2, 3.

68. *Cobalt*, $Co''^{, iv}$, 58.7. *Solution*, CoN_2O_6.

1. $(NH_4)_2S$ precipitates black CoS, insoluble in excess of reagent and in HCl, soluble in aqua regia.

2. KHO or NaHO precipitates blue basic salts, which turn green on exposure to air by oxidation.

3. $(NH_4)HO$ precipitates the same as above, soluble in excess, yielding a reddish-brown fluid; reprecipitated by NaHO or KHO. Acid solutions, or those containing salts of ammonia, are not precipitated.

4. KCN precipitates light-brown $Co(CN)_2$, soluble in excess of the reagent by formation of $2KCN, Co(CN)_2$. This solution is reprecipitated by addition of HCl or H_2SO_4. If to the solution in excess of KCN a few drops of HCl be added, and the solution boiled for some time, $K_3Co(CN)_6$, potassium cobaltic cyanide, is formed, which is not reprecipitated by HCl or H_2SO_4, nor by NaClO.

5. KNO_2 added to cobalt solutions with addition of acetic acid precipitate, on standing, a yellow crystalline double salt.

6. Compounds of cobalt fused with borax in either blow-pipe flame yield deep-blue beads.

69. *Zinc*, Zn'', 65.2. *Solution*, $ZnSO_4$.

1. $(NH_4)_2S$ precipitates white ZnS, insoluble in excess of reagent and in KHO; soluble in the mineral acids.

2. KHO or NaHO precipitates white $Zn(HO)_2$, soluble in excess of either reagent and in $(NH_4)HO$; reprecipitated from dilute solutions by boiling, but not by NH_4Cl.

3. Na_2CO_3 precipitates white basic carbonate, insoluble in excess of reagent.

4. $(NH_4)_2CO_3$ precipitates also the basic carbonate, but it is soluble in excess of the reagent.

5. Heated on charcoal with Na_2CO_3 in the reducing blow-pipe flame, a yellow incrustation of ZnO is obtained, which becomes white when cold.

6. Heated on charcoal by the blow-pipe flame, after moistening with $CoCl_2$ solution, an infusible green mass is obtained.

Characteristic reaction, 1.

70. *Aluminium*, Al^{iv}, 27.3. *Solution*, $Al_2.3SO_4$.

1. $(NH_4)_2S$ precipitates white $Al_2(HO)_6$, soluble in acids.

2. KHO or NaHO produces also a precipitate of $Al_2(HO)_6$, soluble in acids, even in hot acetic acid, and in excess of the reagent. This solution is not precipitated by H_2S, but is reprecipitated by NH_4Cl, or by adding $(NH_4)HO$ after acidifying with HCl.

3. $(NH_4)HO$ also precipitates $Al_2(HO)_6$, soluble in a very large excess of the reagent, more difficultly soluble in presence of salts of ammonia.

4. $BaCO_3$ produces a precipitate of $Al_2(HO)_6$ mixed with basic salt.

5. Na_2HPO_4 precipitates aluminium phosphate, insoluble in $(NH_4)HO$ and in NH_4Cl, but soluble in KHO or NaHO, and in acids. It does not, however, dissolve in hot acetic acid like aluminium hydrate.

6. Heated on charcoal in the blow-pipe flame, then moistened with $CoCl_2$, and reheated, an infusible blue mass is obtained.

Characteristic reactions, 2, 6.

71. *Manganese*, $Mn^{ii, iv}$, 55. *Solution*, $MnSO_4$.

1. $(NH_4)_2S$ produces a flesh-colored precipitate of MnS, soluble in acids, even in acetic acid.

2. KHO or NaHO produces a dirty-white precipitate of $Mn(HO)_2$, insoluble in excess of the reagent; the precipitate rapidly darkens in color by absorption of oxygen. The freshly-precipitated hydrate is dissolved by NH_4Cl, but the higher oxide is insoluble.

3. $(NH_4)HO$ produces the same precipitate of $Mn(HO)_2$, insoluble in excess of the reagent; but it gives no precipitate if the manganese solution contain NH_4Cl. Such a solution on standing precipitates the dark-brown hydrate.

4. Na_2CO_3 produces a white precipitate of $MnCO_3$, which darkens in color by absorption of oxygen.*

5. If any manganese solution (free from chlorine) be treated with PbO_2 and then boiled with HNO_3, it is converted into permanganate, which is recognized by its pink color as soon as the mixture has settled.

6. If any manganese compound be fused on platinum foil with Na_2CO_3 and a trace of KNO_3, it is converted into Na_2MnO_4, recognized by its bright-green color.

7. Fused with borax in the oxidizing flame, an amethyst colored bead is obtained, which becomes colorless in the reducing flame.

Characteristic reactions, 1, 6, 7.

72. *Chromium*, Cr^{iv}, 52.1. *Solution*, $Cr_2 . 3SO_4$.

1. $(NH_4)_2S$ produces a bluish-green precipitate of $Cr_2(HO)_6$, insoluble in excess of the reagent, soluble in acids.

2. $(NH_4)HO$ also precipitates the hydrate, soluble to some extent in excess, yielding a pink fluid, but on heating, the precipitation is complete.

3. KHO or NaHO precipitates also $Cr_2(HO)_6$, soluble

however in excess, yielding a green or bluish-violet solution. On continued boiling or addition of NH_4Cl and heating, the hydrate is reprecipitated.

4. $BaCO_3$ produces a precipitate of $Cr_2(HO)_6$ along with basic salt; the precipitation is not complete till the mixture has stood some time.

5. Fused with Na_2CO_3 and KNO_3 on platinum foil, yellow Na_2CrO_4 is obtained.

6. Fused with borax in either flame (but best in the reducing flame), green beads are obtained.

Characteristic reactions, color of solutions and bead.

TABLE

73. Iron Group (III.). *Separation of Iron, Nickel,*

To filtrate from the sulphides of the Cu and As groups add $(NH_4)HO$ (till shake for some time. Filter. Wash well with H_2O, containing $(NH_4)_2S$, dilute HCl,

RESIDUE.

NiS and CoS.

Test for Co by borax bead. Dissolve the black residue in HCl and $KClO_3$. Boil down just to dryness, dilute with H_2O, add KCN in excess, then a drop of acetic acid, boil for a few minutes, add NaClO in excess, and boil again. A black precipitate indicates Nickel. The filtrate from this precipitate may be tested for Co by evaporating to dryness, and fusing in a borax bead. Blue color indicates Cobalt.

METHOD I. Cr is absent.

Boil down with a little $KClO_3$ till it smells of Cl. Add pure NaHO till strongly alkaline. Filter.

RESIDUE.		FILTRATE.
$Fe_2(HO)_6$, $Mn(HO)_2$. Wash with hot H_2O, dissolve in HCl, add $(NH_4)HO$, and filter.		Al, Zn. Divide into two parts.
RESIDUE.	FILTRATE.	1. Add H_2S or $(NH_4)_2S$. A white precipitate indicates Zinc. Confirm by flame reaction.
$Fe_2(HO)_6$. Dissolve in HCl. Test with $K_4Fe(CN)_6$. Blue precipitate indicates Iron. To ascertain whether the iron is present as ferrous or ferric salt, the *original solution* must be tested with $K_4Fe(CN)_6$, and $K_3Fe(CN)_6$.	Mn. Boil down and ignite, to expel salts of ammonium. Fuse with NaHO and KNO_3. A green residue indicates Manganese. Traces of Ni and Co are found along with the Mn.	2. Add HCl till acid, then $(NH_4)HO$ till alkaline. A white precipitate indicates Aluminium. Confirm by flame reaction.

V.

Cobalt, Aluminium, Zinc, Manganese, and Chromium.

alkaline) + NH_4Cl + $(NH_4)_2S$. Warm the mixture gently in a small flask and and finally once with H_2O alone. Treat the precipitate with cold and filter.

FILTRATE.

Cr, Al, Fe, Zn, and Mn.

(Green or violet if Cr be present. Boil down a portion, and test for Cr by borax bead. Adopt Method I. if absent, Method II. if present.)

METHOD II. Cr is present.

Boil down with a little $KClO_3$ till it smells of Cl. Add Na_2CO_3 or NaHO till just neutral or slightly acid; allow to become *perfectly* cold. Add excess of $BaCO_3$, place in a flask, cork up and shake well, allow to stand till clear. Filter.

RESIDUE.

$Fe_2(HO)_6$, $Cr_2(HO)_6$, $Al_2(HO)_6$ (also excess of $BaCO_3$).

Wash well, boil with pure NaHO, and filter; add HCl to the filtrate, and then $(NH_4)HO$ till alkaline. A white precipitate indicates Aluminium. Confirm by flame reaction. Fuse the residue insoluble in NaHO with a mixture of Na_2CO_3 and KNO_3; extract with water, and filter.

RESIDUE.

$Fe_2(HO)_6$.

Dissolve in HCl, and test with $K_4Fe(CN)_6$. A blue precipitate indicates Iron.

FILTRATE.

Cr.

Yellow in color. Acidify with acetic acid; add lead acetate. A bright-yellow precipitate indicates Chromium.

FILTRATE.

Zn, Mn.

Precipitate the Ba present with H_2SO_4 in the hot solution. Boil well, and filter; add NaHO. Precipitate indicates Manganese. Confirm by fusing with Na_2CO_3 and KNO_3 on platinum foil. To the filtrate from the $Mn(HO)_2$ add $(NH_4)_2S$. A white precipitate indicates Zinc. Confirm by flame reaction.

Reactions of the Metals of the Barium Group (Group IV.).

74. Metals whose carbonates are insoluble in water, and whose solutions are precipitated on the addition of $(NH_4)_2CO_3$. As, however, the carbonates are soluble in acids, the solution, if acid, must be neutralized by addition of $(NH_4)HO$.

BARIUM, STRONTIUM, CALCIUM.

Barium, Ba'', 137. *Solution*, $BaCl_2$.

1. $(NH_4)_2CO_3$ precipitates white $BaCO_3$, soluble in acids, and to a slight extent in NH_4Cl.
2. K_2CO_3 or Na_2CO_3 precipitates also $BaCO_3$, insoluble in excess of either reagent.
3. H_2SO_4 or any soluble sulphate, even in dilute solutions, precipitates heavy white $BaSO_4$, insoluble in acids, alkalies, or salts of ammonium.
4. $CaSO_4$ or $SrSO_4$ precipitates immediately white $BaSO_4$.
5. H_2SiF_6 precipitates white $BaSiF_6$.
6. $C_2(NH_4)_2O_4$ precipitates white C_2BaO_4, soluble in HCl and in HNO_3.
7. K_2CrO_4 precipitates yellow $BaCrO_4$, insoluble in $C_2H_4O_2$, but soluble in HCl and HNO_3.
8. Heated in the lamp flame, a green coloration is produced, especially on moistening the salt with HCl.

Characteristic reactions, 3, 8.

75. *Strontium*, Sr'', 87.5. *Solution*, $SrCl_2$.

1. $(NH_4)_2CO_3$ or K_2CO_3 precipitates white $SrCO_3$, soluble in acids, but less soluble in NH_4Cl than $BaCO_3$.

2. H_2SO_4 precipitates white $SrSO_4$, much less insoluble in H_2O than $BaSO_4$; it therefore precipitates from dilute solutions only on standing or warming. $SrSO_4$ is slightly soluble in HCl.

3. $CaSO_4$, after standing some time, precipitates white $SrSO_4$.

4. H_2SiF_6 does not precipitate strontium solutions.

5. $C_2(NH_4)_2O_4$ precipitates white C_2SrO_4, soluble in HCl and in HNO_3, also to a slight extent in NH_4Cl, but very sparingly in $C_2H_4O_2$.

6. K_2CrO_4, only in concentrated solutions, precipitates yellow $SrCrO_4$, soluble in $C_2H_4O_2$.

7. Heated in the lamp flame, a crimson coloration is produced, especially on moistening the salt with HCl.

Characteristic reactions, 3, 7.

76. *Calcium*, Ca'', 40. *Solution*, $CaCl_2$.

1. $(NH_4)_2CO_3$ or K_2CO_3 precipitates white $CaCO_3$, which becomes crystalline on heating.

2. H_2SO_4 precipitates from strong solutions of calcium salts $CaSO_4$ as a white precipitate, which dissolves in a large excess of water, and also in acids.

3. $CaSO_4$ produces no precipitate.

4. H_2SiF_6 produces no precipitate.

5. $C_2(NH_4)_2O_4$, even in dilute solutions, precipitates white C_2CaO_4, soluble in HCl or HNO_3, but insoluble in $C_2H_2O_4$ or in $C_2H_4O_2$.

6. Heated in the lamp flame, a dull-red coloration is produced, especially on moistening the salt with HCl. This reaction is imperceptible in presence of Ba or Sr salts.

Characteristic reactions, 5, 6.

TABLE VI.

77. Barium Group (IV.). *Separation of Barium, Strontium, and Calcium.*

Heat filtrate from iron group, add to the hot solution NH_4Cl and $(NH_4)_2CO_3$, and filter. Wash precipitate with hot H_2O, dissolve in HCl, and add $CaSO_4$ solution. An immediate precipitate indicates **Barium**; a precipitate after some time indicates Strontium, or a dilute Barium solution. (Test another portion with $SrSO_4$ for Ba.) To another portion of the solution in HCl add H_2SO_4, and boil to remove Ba and Sr. Filter. Neutralize filtrate with $(NH_4)HO$, and add $C_2(NH_4)_2O_4$. An immediate precipitate indicates Calcium.

TO TEST FOR STRONTIUM.

I. Ba present; Ca absent.

Dissolve the carbonate in HCl, and evaporate to dryness. Treat the residue with strong alcohol. Filter.

RESIDUE.	FILTRATE.
$BaCl_2$.	Sr.
Confirm by flame test. Green coloration indicates Barium.	Confirm by lighting the alcoholic solution. Crimson coloration indicates Strontium.

II. Ba absent; Ca present.

Dissolve as before, and precipitate with H_2SO_4. Filter, and wash well.

RESIDUE.	FILTRATE.
$SrSO_4$.	Ca.
If small, burn the filter in the reducing gas flame to convert $SrSO_4$ into SrS; moisten with HCl, and test in the lamp flame. Crimson coloration indicates Strontium.	Neutralize the solution with $(NH_4)HO$, and test with $C_2(NH_4)_2O_4$. White precipitate indicates Calcium.

Another Method.

Dissolve in HNO_3, and evaporate to dryness. Treat with strong alcohol. Filter.

Residue.	Filtrate.
$Sr(NO_3)_2$. Confirm as above.	Ca. Confirm as above.

III. Ba and Ca present.

Add H_2SO_4 to the HCl solution (diluted to prevent precipitation of Calcium), and filter.

Residue.	Filtrate.
$BaSO_4$, $SrSO_4$. Boil in a beaker with a little water, together with a mixture of three parts K_2SO_4 and one part K_2CO_3. Filter, and treat residue with HNO_3. The $SrSO_4$ is dissolved, and the $BaSO_4$ left undissolved. (Traces of Ca may be found with the Sr.)	Ca. Neutralize the solution with $(NH_4)HO$, and test with $C_2(NH_4)_2O_4$. White precipitate indicates Calcium.

Another Method.

Dissolve the carbonates in $C_2H_4O_2$, and precipitate the Ba with K_2CrO_4. Filter. Precipitate the Sr and Ca by $(NH_4)_2CO_3$, and proceed as in Method II. (Ba absent, Ca present).

Reactions of the Metals of the Potassium Group (Group V.).

78. Metals whose solutions are unprecipitated by the preceding group reagents, and having no common precipitant. They are therefore recognized by individual tests.

MAGNESIUM, POTASSIUM, AMMONIUM, SODIUM.

Magnesium, Mg'', 24. *Solution*, $MgSO_4$.

1. $(NH_4)HO$ and $(NH_4)_2CO_3$ give no precipitates in the presence of salts of ammonium.

2. Na_2HPO_4 in presence of $(NH_4)HO$ and $(NH_4)Cl$ a crystalline white precipitate of $MgNH_4PO_4$. The precipitation is slow from dilute solutions, but may be hastened by stirring with a glass rod and warming. Precipitate soluble in dilute mineral acids and in $C_2H_4O_2$, almost insoluble in dilute solution of $(NH_4)HO$.

3. H_2SO_4, H_2SiF_6, and $C_2(NH_4)_2O_4$ give no precipitates.

4. Heated on charcoal in the blow-pipe flame, and then moistened with $CoCl_2$ or CoN_2O_6 and reheated, gives a pink mass.

Characteristic reactions, 2, 4.

79. *Potassium*, K', 39.1. *Solution*, KCl.

1. $PtCl_4$, except in dilute solutions, gives a crystalline yellow precipitate of $2KCl + PtCl_4$. The precipitation is hastened by stirring or the addition of alcohol.

2. H_2Tr or NaHTr precipitates white crystalline KHTr from concentrated solutions.

3. H_2SiF_6 precipitates white gelatinous K_2SiF_6.

4. Heated on platinum wire, potassium compounds color the flame violet, appearing reddish-violet through blue glass.

Characteristic reaction, 4.

80. *Ammonium*, NH_4, 18. *Solution*, NH_4Cl.

1. $PtCl_4$ gives a crystalline yellow precipitate of $2NH_4Cl + PtCl_4$, except in dilute solutions. Precipitate insoluble in alcohol and ether. On ignition, precipitate leaves a residue of spongy platinum.

2. NaHTr or H_2Tr give in strong solutions a white precipitate of $(NH_4)HTr$.

3. H_2SiF_6 gives no precipitate.

4. Nessler's solution gives a brown precipitate, or in any dilute solutions a yellow coloration.

5. Heated with NaHO or KHO, compounds of ammonium evolve ammonia gas, recognized by its odor, alkaline reaction, and fuming with HCl.

6. Heated on platinum foil, all compounds of ammonia volatilize completely.

Characteristic reactions, 4, 5, 6.

81. *Sodium*, Na', 23. *Solution*, NaCl.

1. $PtCl_4$, NaHTr, and H_2Tr give no precipitates.

2. H_2SiF_6 gives precipitate of white gelatinous Na_2SiF_6.

3. The salts of sodium being almost without an exception soluble in water, the flame test alone serves to distinguish the metal. Heated on platinum foil or wire in non-luminous flame, an intense yellow color is produced, not seen when viewed through blue glass.

TABLE VII.

82. Group V. *Separation of Magnesium, Potassium, Sodium, and Ammonium.*

The filtrate from the Barium group is concentrated by evaporation and a portion ignited on platinum foil. If no residue is left on ignition, Mg, K, and Na are absent.

Detection of NH_4.

The original substance or solution is heated with NaHO in a test-tube. Presence of Ammonium shown by smell, by the white fumes with HCl, and by its action on red litmus paper.

Detection of Mg.

To a portion of the concentrated cold solution add $(NH_4)HO$ and Na_2HPO_4. White crystalline precipitate denotes Magnesium.

Detection of K and Na.

(1.) Mg being absent.

Evaporate another portion of the solution to dryness, ignite residue, dissolve in a small quantity of water, filter if required, and add to the clear liquid $PtCl_4$, evaporate nearly to dryness, and add alcohol. Yellow precipitate indicates Potassium.

(2.) Mg being present.

Evaporate the solution to dryness, ignite residue, dissolve in water, and add baryta water until the solution has an alkaline reaction; boil; filter. To filtrate, add $(NH_4)_2CO_3$, heat, filter, evaporate to dryness, and test the residue for K and Na.

To detect Na.

Evaporate alcoholic solution (which must have a yellow color, showing that excess of $PtCl_4$ has been added) nearly to dryness, add a grain or two of sugar, and ignite residue. Exhaust with water, filter, evaporate to dryness; and if a residue be left, test it by flame reaction for Na. Yellow coloration indicates Sodium.

PRECIPITATIONS IN THE FIVE GROUPS OF BASES.

Precipitated by	GROUP I.	GROUP II.	GROUP III.	GROUP IV.	GROUP V.
I. Chlorides . . .	Pb,† Ag, Hg_2.				
II. H_2S in Acid Sol.	Pb, Ag, Hg_2.	As, Sb, Sn, Bi, Cu, Cd, Hg.			
III. Sulphides . . .	Pb, Ag, Hg_2.	As,* Sb,* Sn,* Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Cr, Co, Ni.		
IV. Carbonates with $(NH_4)Cl$	Pb, Ag, Hg_2.	Sb,* Sn, Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Cr, Co, Ni.	Ba, Sr, Ca.	
Carbonates	Pb, Ag, Hg_2.	Sb,* Sn, Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Cr, Co, Ni.	Ba, Sr, Ca.	Mg.
Fixed Alkalies . . .	Pb,* Ag, Hg_2.	Sb,* Sn,* Bi, Cu, Cd, Hg.	Zn,* Al,* Fe, Mn, Cr,* Co, Ni.	Ba,† Sr,† Ca.	Mg.
Ammonia	Pb, Ag,* Hg_2.	Sb, Sn, Bi, Cu,* Cd,* Hg.	Zn,* Al, Fe, Mn, Cr, Co,* Ni.*		Mg.
Phosphates	Pb, Ag, Hg_2.	Sb, Sn, Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Cr, Co, Ni.	Ba, Sr, Ca.	Mg.
Oxalates	Pb, Ag, Hg_2.	Sb, Sn, Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Co, Ni.	Ba,† Sr, Ca.	Mg.†
Sulphates	Pb.			Ba, Sr, Ca.†	
Sulphites	Pb, Ag, Hg_2.	Sb, Sn, Bi, Cu, Cd, Hg.	Zn, Al, Fe, Mn, Co, Ni.	Ba, Sr, Ca.	Mg.
Bromides	Pb, Ag, Hg_2.				
Iodides	Pb,* Ag,* Hg.	(Bi), Hg.*			

* Soluble in excess of reagent. † Not precipitated in dilute solutions.

Reactions of the Acids.

83. Grouping of the Acids. The acids can be approximately classified by means of certain group reagents. They are divided into two great classes: *inorganic* and *organic acids*. These are easily distinguished by the action of heat.

Salts of *inorganic acids*, when heated to redness, are not charred; salts of *organic acids* are at once charred, owing to decomposition and separation of carbon (with the exception of acetic and formic acids).

84. Grouping of the Inorganic Acids.

Group I. (*Sulphuric Acid Group.*)

Group reagent, $BaCl_2$ in presence of HCl.

Sulphuric acid, hydrofluo-silicic acid.

The acids of this group are precipitated by $BaCl_2$, and the precipitate is not dissolved on addition of HCl.

Group II. (*Phosphoric Acid Group.*)

Group reagent, $BaCl_2$.

Phosphoric, boric, hydrofluoric, carbonic, silicic, sulphurous, arsenious, arsenic, iodic, chromic acids.

The acids of this group are precipitated in *neutral solutions* by $BaCl_2$.

Group III. (*Hydrochloric Acid Group.*)

Group reagent, $AgNO_3$.

Hydrochloric, hydrobromic, hydroiodic, hydrocyanic, and hydrosulphuric acids.

The acids of this group are precipitated by $AgNO_3$, and not by $BaCl_2$.

Group IV. (*Nitric Acid Group.*)

Nitric, chloric, and perchloric acids.

These acids are not precipitated by any reagent, as all their salts are soluble in water.

Reactions of the Inorganic Acids belonging to Group I.

85. Acids precipitated by $BaCl_2$ in presence of HCl.

SULPHURIC ACID, HYDROFLUO-SILICIC ACID.

Sulphuric Acid, H_2SO_4, 98.

1. $BaCl_2$ precipitates a white $BaSO_4$, insoluble in HCl or HNO_3. In very dilute solutions the precipitation is not immediate, but on standing, the solution becomes clouded, and ultimately the precipitate subsides.

2. $Pb(NO_3)_2$ precipitates a heavy white $PbSO_4$, soluble in NaHO, and in boiling HCl (on allowing this solution to cool, $PbCl_2$ crystallizes out).

3. Fused on charcoal with Na_2CO_3 in the reducing flame of the blow-pipe, a sulphide is produced. If the fused mass be moistened with HCl, the odor of H_2S is at once perceptible; or if it be placed on a bright piece of silver and moistened with water, a black stain of Ag_2S is produced.

86. *Hydrofluo-silicic Acid*, H_2SiF_6, 144.

1. $BaCl_2$ precipitates a crystalline $BaSiF_6$, insoluble in HCl.

2. KCl precipitates a gelatinous K_2SiF_6.

3. Heated with H_2SO_4 in a leaden crucible covered with

a piece of glass, the latter will be etched by the evolved HF.

Reactions of the Acids belonging to Group II.

87. Acids precipitated by $BaCl_2$ in neutral solutions.

PHOSPHORIC, BORIC, HYDROFLUORIC, CARBONIC, SILICIC, SULPHUROUS, ARSENIOUS, ARSENIC, IODIC, AND CHROMIC ACIDS.

Phosphoric Acid, H_3PO_4, 98.

1. $BaCl_2$ precipitates a white $BaHPO_4$, readily soluble in HNO_3 or HCl, but with difficulty in NH_4Cl.

2. Mg_2SO_4, along with $(NH_4)HO$ and NH_4Cl, precipitates a white crystalline $Mg(NH_4)PO_4 + 6H_2O$, insoluble in $(NH_4)HO$, but soluble in HCl, HNO_3, and acetic acid. In dilute solutions the precipitation does not take place till after the lapse of some time, but is promoted by stirring and gentle warming.

3. $AgNO_3$ precipitates a yellow Ag_3PO_4, soluble in HNO_3, and also in $(NH_4)HO$.

4. Lead acetate precipitates a white $Pb_3(PO_4)_2$, soluble in HNO_3, but almost insoluble in acetic acid.

5. Fe_2Cl_6, in presence of excess of sodium acetate, precipitates a yellowish $FePO_4$, soluble in HCl, and in excess of Fe_2Cl_6, which must be added drop by drop.

6. Ammonium molybdate produces in solutions acidified by HNO_3 a yellow color, and then a precipitate; this reaction is hastened by warming.

88. *Boric Acid*, $B(HO)_3$, 62.

1. $BaCl_2$ precipitates a white $Pa(BO_2)_2$, soluble in acids.

2. $AgNO_3$ produces in strong solutions a yellowish-white precipitate. In dilute solutions Ag_2O is precipitated.

3. H_2SO_4 or HCl, added to hot concentrated solutions of alkaline borates, on cooling, precipitates a crystalline $B(HO)_3$.

4. If alcohol containing free boric acid be kindled, it burns with a green flame, best seen on stirring the mixture. Borates may be examined in this way by first adding strong H_2SO_4, to liberate the $B(HO)_3$.

5. If the solution of a borate be made distinctly acid with HCl, and turmeric paper dipped into it, the latter, on gentle warming, acquires a brown tint, which is turned blue by caustic soda.

89. *Hydrofluoric Acid*, HF, 20.

1. $BaCl_2$ precipitates a white BaF_2, soluble in HCl, and sparingly in NH_4Cl.

2. $CaCl_2$ produces a gelatinous and almost transparent precipitate of CaF_2, made more apparent on addition of $(NH_4)HO$. The precipitate is very difficultly soluble in HCl, even on boiling, and is nearly insoluble in acetic acid.

3. Heated with H_2SO_4, all fluorides are decomposed with evolution of HF, which is recognized by its power of etching glass.

4. Heated with a mixture of borax and $HKSO_4$, on a loop of platinum wire in the non-luminous flame, BF_3 is produced, coloring the flame green.

90. *Carbonic Acid*, H_2CO_3, $H_2O + CO_2$.

1. $BaCl_2$ in neutral solutions precipitates a white $BaCO_3$, soluble in acids with effervescence.

2. Treated with dilute HCl, all carbonates at once evolve CO_2 with effervescence, which turns lime water a milky white from the formation of $CaCO_3$.

91. *Silicic Acid*, $Si(HO)_4$, 96.

1. $BaCl_2$ precipitates a white $SiBa_2O_4$, which is decomposed on addition of HCl, and $Si(HO)_4$ separates out as a gelatinous precipitate.

2. HCl, added drop by drop to a strong solution of a silicate, precipitates a gelatinous $Si(HO)_4$; but if added to a dilute solution or in large excess, no precipitate is obtained until the mixture has been evaporated to dryness and ignited, when SiO_2 separates out, and this is not redissolved on addition of HCl.

3. Fused with Na_2CO_3 in a loop of platinum wire in the non-luminous gas-flame, effervescence occurs from the disengagement of CO_2, and the bead is transparent on cooling, unless the Na_2CO_3 be in excess.

4. Fused with microcosmic salt on a loop of platinum wire in the non-luminous gas-flame, solution does not take place, but the silica floats about on the bead undissolved.

92. *Sulphurous Acid*, H_2SO_3, 82.

1. $BaCl_2$ precipitates a white $BaSO_3$, soluble in HCl. On addition of chlorine water, gives a white precipitate of $BaSO_4$, the sulphite being oxidized to the sulphate.

2. $AgNO_3$ gives a white precipitate of $AgSO_3$, darkened on heating.

3. Added to a mixture of Zn and HCl, H_2S is produced, and recognized by its smell and by its action on paper moistened with a solution of a lead salt, blackening it.

4. H_2S decomposes pure H_2SO_3, with separation of sulphur.

5. H_2SO_3 is decomposed by HCl, with evolution of SO_2.

93. *Arsenious Acid*, H_3AsO_3, 126.

1. $AgNO_3$ gives in neutral solutions a yellow precipitate of Ag_3AsO_3, soluble in $(NH_4)HO$.

2. $MgSO_4 + (NH_4)Cl + (NH_4)HO$ give no precipitate.

3. H_2S precipitates As_2S_3, yellow.

94. *Arsenic Acid*, H_3AsO_4, 142.

1. $AgNO_3$ gives in neutral solutions a light-brown precipitate of $AgAsO_4$.

2. $MgSO_4 + (NH_4)Cl + (NH_4)HO$ give a white precipitate of $MgNH_4AsO_4$.

3. H_2S precipitates As_2S_3, yellow.

95. *Iodic Acid*, HIO_3, 176.

1. $BaCl_2$ gives a white precipitate of BaI_2O_6, soluble in HNO_3.

2. $AgNO_3$ precipitates white crystalline $AgIO_3$, easily soluble in $(NH_4)HO$, but sparingly so in HNO_3.

3. SO_2 gives at first a precipitate of I, which is converted into HI on addition of excess of reagent.

4. HIO_3 is decomposed by H_2S, with formation of an iodide and separation of S.

5. Iodate salts, on heating, are decomposed, oxygen being evolved. In some cases iodine is given off in violet vapors.

96. *Chromic Acid*, H_2CrO_4, 118.2.

1. $BaCl_2$ precipitates a yellow $BaCrO_4$, soluble in HCl and HNO_3, but insoluble in acetic acid.

2. H_2S in presence of HCl reduces the solution to Cr_2Cl_6 (green), with separation of S. In neutral solutions $Cr_2(HO)$ is precipitated along with S.

3. SO_2 reduces solutions of chromates to the chromic salt, the color of which is green. Chromates are likewise reduced by zinc and a dilute acid, by oxalic acid and dilute sulphuric acid, by strong H_2SO_4, by strong HCl, and by boiling the solution, acidified with HCl or H_2SO_4, along with alcohol.

4. $AgNO_3$ precipitates a dark-red Ag_2CrO_4, soluble in HNO_3 and in $(NH_4)HO$.

5. Lead acetate produces a bright-yellow precipitate of $PbCrO_4$, soluble in NaHO, but soluble with difficulty in dilute HNO_3.

6. H_2CrO_4 is precipitated by $(NH_4)_2S$ as $Cr_2(HO)_6$.

Reactions of the Acids belonging to Group III.

97. Acids precipitated by $AgNO_3$, and not by $BaCl_2$.

HYDROCHLORIC, HYDROBROMIC, HYDRIODIC, HYDROCYANIC, AND HYDROSULPHURIC ACIDS.

Hydrochloric Acid, HCl, 36.5.

1. $AgNO_3$ precipitates a white curdy AgCl, which becomes violet on exposure to light. The precipitate is insoluble in HNO_3, but soluble in $(NH_4)HO$, in KCN, in $Na_2S_2O_3$, and also to some extent in NaCl.

2. Heated with H_2SO_4 and MnO_2, chlorides yield chlorine gas, recognized by its smell, bleaching action, and green color.

3. *Dry* chlorides, when heated in a retort with H_2SO_4 and $K_2Cr_2O_7$, yield CrO_2Cl_2 (chromium oxychloride),

which distils over into the receiver as a dark-red liquid, decomposed by addition of water or $(NH_4)HO$, yielding a yellow solution, which, on addition of a lead salt, gives a yellow precipitate of $PbCrO_4$.

98. *Hydrobromic Acid*, HBr, 81.

1. $AgNO_3$ precipitates a pale-yellow AgBr, insoluble in dilute HNO_3, soluble in strong $(NH_4)HO$, and readily in KCN and $Na_2S_2O_3$.

2. Heated with H_2SO_4 and MnO_2, bromides yield red vapors of Br, recognized by its powerful odor.

3. Heated in a retort with $K_2Cr_2O_7$ and H_2SO_4, dry bromides yield dark-red vapors, which condense in the receiver to a liquid of the same color, which consists of pure bromine, and is decolorized on adding excess of $(NH_4)HO$.

99. *Hydriodic Acid*, HI, 128.

1. $AgNO_3$ precipitates a pale-yellow AgI, insoluble in dilute HNO_3, and very difficultly soluble in $(NH_4)HO$, but readily in KCN and $Na_2S_2O_3$.

2. Cuprous sulphate precipitates a dirty-white Cu_2I_2, which separates most completely if the solution be made slightly alkaline with Na_2CO_3.

3. KNO_2 produces no reaction in solutions of iodides until a few drops of HCl or H_2SO_4 are added, when iodine is at once liberated and colors the solution yellow. If a little starch solution be now added, a deep-blue coloration results from the formation of starch iodide.

4. Chlorine water (or the gas) liberates iodine from iodides, but *excess* of Cl causes the formation of ICl_3,

which is colorless, and gives no blue coloration with starch solution.

5. Heated with MnO_2 and dilute H_2SO_4, violet vapors of iodine are obtained, which color starch paper blue.

100. *Hydrocyanic Acid*, HCN, 27.

1. $AgNO_3$ precipitates a white AgCN, insoluble in HNO_3, with difficulty in $(NH_4)HO$, but readily in KCN and $Na_2S_2O_3$. AgCN is decomposed on ignition, and metallic Ag remains; this serves to distinguish it from AgCl, which is not decomposed on ignition.

2. If a solution of $FeSO_4$ which has become oxidized by exposure to the air, be added to the solution of a cyanide made alkaline with NaHO, a bluish-green precipitate is formed, which is a mixture of Prussian blue with the hydrated oxides of iron. On adding HCl, these last are dissolved, and the blue precipitate remains.

3. HCl decomposes nearly all cyanides with evolution of HCN, recognized by its odor, resembling bitter almonds.

4. $Hg(CN)_2$ cannot be detected by the above methods. The dry substance is detected by igniting in a small tube, when cyanogen gas is evolved, or the solution is decomposed by H_2S and filtered from the HgS: the filtrate contains HCN.

101. *Hydrosulphuric Acid (Sulphuretted Hydrogen)*, H_2S, 34.

1. $AgNO_3$ precipitates a black Ag_2S, insoluble in dilute acids.

2. Lead acetate, even when highly dilute, precipitates black PbS.

3. Sodium nitro-prusside, in presence of NaHO, produces a reddish-violet coloration, even in very dilute solutions. The color disappears in a short time.

4. HCl or H_2SO_4 decomposes most sulphides with evolution of H_2S, recognized by its disagreeable odor and by its blackening paper moistened with solution of lead.

Reactions of the Acids of Group IV.

102. Acids not precipitated by any reagent.

Nitric Acid, HNO_3, 63.

1. Nitrates when heated evolve oxygen, and in some cases nitrous vapors also. On fusing a nitrate and adding a fragment of charcoal, vivid deflagration occurs.

2. Free HNO_3 heated with Cu gives red fumes; boiled with pieces of silk or wool turns them yellow.

3. If to a solution of a nitrate $FeSO_4$ and concentrated H_2SO_4 be poured carefully into the test-tube, a dark ring will appear on top of the H_2SO_4, which will be violet, red, or dark-brown according to the quantity of HNO_3 present. The ring disappears on warming.

103. *Chloric Acid*, $HClO_3$, 84.5.

1. H_2SO_4 decomposes chlorates with evolution of Cl_2O_4, a greenish-yellow gas having a powerful odor. If heated, violent explosions occur; the mixture ought therefore to be kept cold, and only *very* small quantities should be used.

2. When chlorates are heated, oxygen is evolved, and a metallic chloride remains, which may be dissolved in water, and precipitated as AgCl by $AgNO_3$.

3. Chlorates are reduced by SO_2 with liberation of chlorine or its oxides; hence if the solution of a chlorate be colored blue with indigo solution, it is decolorized on adding H_2SO_4 and solution of Na_2SO_3. (Distinction from perchlorates.)

4. HCl decomposes chlorates with evolution of Cl and Cl_2O_4, a mixture called *euchlorine*.

5. Heated with charcoal, chlorates deflagrate violently.

104. *Perchloric Acid*, $HClO_4$, 100.5.

1. H_2SO_4 does not act upon perchlorates in the cold, and on heating, white fumes of $HClO_4$ are given off, but no explosions occur.

2. KCl in strong solutions precipitates a white $KClO_4$.

3. Indigo solution is not decolorized when added to perchlorates warmed with HCl, as euchlorine is not evolved.

4. Dry perchlorates evolve oxygen on heating.

5. Perchlorates are not reduced by SO_2.

TABLE VIII.

Detection of Inorganic Acids in Mixtures.

(The following acids are found in examination for bases, which ought *always* to *precede* examination for acids:—

H_2SO_3,
$H_2S_2O_3$,
H_2CO_3,
H_2S.
$Si(HO)_4$,
H_2CrO_4,
H_3AsO_3,
H_3AsO_4.)

I.

Acids in Soluble Bodies.

1. Neutralize a portion of the solution with $(NH_4)HO$, and add $BaCl_2$ (or $Ba(NO_3)_2$ if Ag, Hg_2, or Pb be present): precipitate indicates H_2SO_4, H_3PO_4, H_3AsO_3, H_3AsO_4, $Si(HO)_4$, H_2CrO_4, and large quantities of $B(HO)_3$ and HF.*

To precipitate, add H_2O, and then HCl: if a precipitate remain, H_2SO_4 was present.

2. To another portion of the neutralized solution add $AgNO_3$: a precipitate indicates one or more of these acids; *i.e.*,

(*a*) HCl, HBr, HI, HCn, $H_4Fe(Cn)_6$, $H_3Fe(Cn)_6$, H_2S.

(*b*) H_3PO_4, H_3AsO_4, H_3AsO_3, H_2CrO_4, $Si(HO)_4$, $B(HO)_3$.*

To the precipitate add cold dilute HNO_3. Acids under (*a*) are insoluble; those under (*b*), soluble.

Detection of Acids under (a).

To a portion of the solution add starch paste and one drop of a solution of N_2O_3 in H_2SO_4. Blue coloration indicates HI. Add chlorine water till the blue color disappears, and shake with chloroform. Reddish brown

* Oxalic, citric, and tartaric acids will also be shown, if present.

color indicates the presence of HBr. HCl is detected in presence of the others by boiling down the solution to dryness and distilling residue with $K_2Cr_2O_7$ and H_2SO_4 (see 97, 3).

Detection of Acids under (*b*).

Test separately for each acid by the methods already given.

Separation of H_3AsO_3, H_3AsO_4, *and* H_3PO_4.

Acidify solution with HCl, add Na_2SO_3, and heat until no smell of SO_2 is given off.

Pass H_2S through the hot solution, filter, and test for H_3PO_4 with ammonium molybdate: yellow precipitate indicates H_3PO_4.

Precipitate another portion with magnesia mixture, and test both precipitate and filtrate for arsenic.

Test for the other acids by the following reactions, given under each acid: —

For HCn, by test 3, 100.
For H_2S, by test 4, 101.
For HNO_3, by tests 2 and 2, 102.
For $HClO_3$, by tests 1 and 2, 103.
For $B(HO)_3$, by tests 4 and 5, 88.
For $Si(HO)_4$, by tests 2 and 4, 96.
For H_2SO_3, by test 3, 92, and smell of SO_2 on adding HCl.
For CO_2, by test 2, 90.

A SHORTER SYSTEM OF QUALITATIVE ANALYSIS.

TESTS FOR BASES.

A.

Add dilute HCl.

1. No prpt; pass to *B*.

2. A prpt; filter, filtrate to *B*; prpt may be

Ag, Hg_2, Pb.

Wash prpt twice with cold water, once with hot water; save only hot washing.

1. To hot washing add dilute H_2SO_4; prpt shows Pb.
2. To prpt on filter add NH_4HO.
   a. Prpt turns black; Hg_2 present.
   b. To ammoniacal filtrate add HNO_3 in excess; prpt shows Ag.

B.

Add H_2S water.

1. No prpt; pass to *C*.

2. A prpt. Saturate with H_2S gas and filter; filtrate to *C*. Prpt may be

Cu, Bi, Cd, Pb; Hg; As, Sb, Sn.

Wash prpt carefully, warm with $(NH_4)_2S$, and filter.

1. Prpt. Boil with HNO_3, and filter.
   a. Black residue shows Hg.

b. Solution. Add dilute H_2SO_4; white prpt shows Pb. Filter, and add NH_4HO; white prpt shows Bi, blue solution shows Cu; filter, and to filtrate add H_2S.

aa. Yellow prpt shows Cd.

bb. Brown prpt shows Cu; filter, dissolve prpt in KCy solution; yellow residue shows Cd.

2. Filtrate. Add dilute HCl.

a. Pure white prpt; no As, Sb, Sn.

b. Colored prpt; boil with $KClO_3$ crystals, put in evolution flask with dilute H_2SO_4 and Zn, tube dipping in $AgNO_3$ solution.

aa. In $AgNO_3$ solution a black prpt shows Sb. Filter, to filtrate add more $AgNO_3$ solution, and then, carefully, very dilute NH_4HO; yellow prpt shows As.

bb. In evolution flask remove Zn, add HCl, and warm, then add $HgCl_2$; white prpt shows Sn.

C.

Add NH_4Cl, NH_4HO, and $(NH_4)_2S$.

1. No prpt; go on to *D*.

2. A prpt. Filter, filtrate to *D*. Prpt may be

Ni, Co; Fe, Al, Cr; Mn, Zn.

Treat prpt with cold H_2O and HCl, and filter.

1. Black residue. Test with borax bead; brown shows Ni, blue shows Co or Co and Ni. In latter case dissolve residue in aqua regia, evaporate excess of acid, add $NaC_2H_3O_2$ and KNO_3 and let stand. Yellow prpt after a while shows Co. Filter, to filtrate add KHO; green prpt shows Ni.

2. Filtrate. Boil with $KClO_3$ crystals, nearly neutralize with Na_2CO_3, and let cool; then add $BaCO_3$ and let stand a while, shaking occasionally; then filter.
   - *a.* Prpt. Dissolve in dilute H_2SO_4, boil and filter; if necessary, boil filtrate and refilter; to filtrate add KHO in excess, boil three minutes and filter.
     - *aa.* Prpt. Fuse part on Pt with KNO_3 and Na_2CO_3, dissolve in H_2O, filter, add $C_2H_4O_2$; to filtrate add $Pb(C_2H_3O_2)_2$; yellow prpt shows Cr. Dissolve part in HCl and add KCfy; blue prpt shows Fe.
     - *bb.* Filtrate. Add HCl and NH_4HO; white prpt shows Al.
   - *b.* Filtrate. Add KHO in excess and filter.
     - *aa.* Filtrate. Add H_2S; white prpt shows Zn.
     - *bb.* Prpt. Fuse on Pt foil with KNO_3 and Na_2CO_3; green shows Mn.

D.

Add $(NH_4)_2CO_3$.

1. No prpt; pass to *E*.
2. A prpt; filter, filtrate to *E*. Prpt may be

Ba, Sr, Ca.

Dissolve prpt on filter in $C_2H_4O_2$; to filtrate add $K_2Cr_2O_7$; yellow prpt shows Ba. Filter, to filtrate add K_2SO_4; prpt shows Sr. Filter, to filtrate add $(NH_4)_2C_2O_4$; prpt shows Ca. If no prpt, add dilute H_2SO_4 and alcohol; prpt shows Ca.

E.

1. To some add HNa_2PO_4; prpt shows Mg.
2. Evaporate rest to dryness, and ignite gently. Residue may be K, Na.

Dissolve in very little dilute HCl.

Examine on Pt wire in flame; purple flame shows K; yellow, Na, or Na and K. In latter case add $PtCl_4$; yellow prpt shows K.

F.

Boil some of original solution with KHO; NH_4 shown by its smell.

TESTS FOR ACIDS.

Each acid must be tested for in a fresh portion of the original solution.

Add $BaCl_2$*; white prpt, insoluble in HNO_3, shows H_2SO_4.

Add $AgNO_3$.

1. White prpt, soluble in NH_4HO and reprecipitated by HNO_3, shows HCl.
2. Yellowish prpt, insoluble in NH_4HO, shows HI.

HCl may also be present. Filter, to filtrate add HNO_3; prpt shows HCl.

Add strong H_2SO_4, then strong, fresh solution of $FeSO_4$, in such a way that the two fluids do not mix; shake then very gently; a brown color at junction shows HNO_3.

Add HCl;†

1. Effervescence and smell of H_2S shows H_2S.
2. Effervescence and no smell shows H_2CO_3.

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* If Hg_2, Ag, or Pb are present, use $BaNO_3$ instead of BaCl.

† These reactions must have been noticed in testing for Bases.

TABLE OF EQUIVALENT VALUES AND USEFUL CONSTANTS.

1 meter	= 39.3708 inches.
1 kilometer	= 0.6214 mile.
1 square meter	= 10.7643 square feet.
1 centimeter	= 0.061 cubic inch.
1 liter	= 61.027 cubic inches = 1.0567 quarts.
1 gram	= 15.4823 grains.
1 kilogram	= 2.2046 pounds.
1 kilogrammeter	= 7.23308 foot-pounds.

1 cubic foot of water = 6.314 gals. at 4° C. weighs 62.425 lbs.
1 gallon of water at 17° C. weighs 10 lbs.
1 ton of water occupies 224 gallons, or 35.88 cubic feet.

THREE FUNDAMENTAL UNITS.

A definite length — the Centimeter.
A definite mass — the Gram.
A definite time — the Second.

Value of g at London, in meters, per second	9.817
Value of g at London, in feet, per second	32.19
Mean density of the earth	5.67
Mass of the earth in tons	6×10^{21}
Normal height of the barometer in mm.	760
Volume in liters of the molecular weight of a gas in grams	22.32
Double density of air (mass, 22.32 liters)	28.872
Mass of a liter of normal hydrogen in grams	0.0896
Percentage of oxygen in air by mass	23.18
Percentage of oxygen in air by volume	20.9
Mass of a liter of water at 4° C. in grams	1000.13
Mass of a cubic foot of water at 4° C. in lbs.	62.425
Latent heat of water	79.25
Latent heat of steam	536
Mechanical equivalent of the unit of heat (Kilog-met)	425.434
Mechanical equivalent of the unit of heat (foot-pounds)	775.5

TABLE SHOWING THE SOLUBILITY IN WATER AND ACIDS OF THE COMMON METALLIC SALTS.

	Acetate.	Arsenate.	Arsenite.	Borate.	Bromide.	Carbonate.	Chlorate.	Chloride.	Chromate.	Cyanide.	Ferricyanide.	Ferrocyanide.	Fluoride.	Iodide.	Nitrate.	Oxalate.	Oxide.	Phosphate.	Silicate.	Sulphate.	Sulphide.
Ag	s	i	i	i	I	i	S	I	i	i	I	I	S	i	S	i	i	i		s	i
Pb	S	i	i	i	sI	i	S	sI	i	i	s	i	s	i	S	i	i·	i	i	I	i
Hg_2	s	i	i		i	i	S	i	i	i	i	i	S	i	S	i	i	i		s	i
Hg	S	i	i		s	i	S	S	s	S			S	i	Si	i	i	i		Si	i
Bi	S	i		i	Si	i	S	Si	i	s		I	s	Si	Si	i	i	i		Si	i
Cu	S	i	i	i	S	i	S	S	s	i	I	I	s	Si	S	i	i	i	i	S	i
Cd	S	i		s	S	i	S	S	i	i	i		s	S	S	i	i	i	i	S	i
Sn''	S	i	i	i	S			Si	i		i	i	S	S	Si	i	i	i		S	i
Sn^{iv}	S	i			S			S				i	S	Si	Si	S	i	i		S	i
Sb	S	i	i		Si			Si	i				S	Si		i	i	i		Si	i
Fe''	S	i	i	i	S	i	S	S		i	I	I	s	S	S	i	i	i	i	S	i
Fe^{iv}	S	i	i	i	S		S	S	s		S	I	s	S	S	s	i	i	i	S	i
Cr	S	i		i	S		S	S	i	i			s	S	S	S	i	i	i	S	i
Al	S	i		i	S		S	S					i		S	i	i	i		S	i
Mn	S	i	i	i	S	i	S	S	s	i	I	i	i	S	S	i	i	i	i	S	i
Zn	S	i		i	S	i	S	S	S	i	i	i	s	S	S	i	i	i	i	S	i
Co	S	i	i	i	S	i	S	S	i	i	I	I	s	S	S	i	i	i	i	S	i
Ni	S	i	i	i	S	i	S	S	i	i	I	I	s	S	S	i	i	i	i	S	i
Ba	S	i	s	i	S	i	S	S	i	s	s	s	i	S	S	i		i	i	I	S
Sr	S	i	s	i	S	i	S	S	s	S		S	i	S	S	i	i	i	i	I	S
Ca	S	i	i	i	S	i	S	S	s	S	S	S	i	S	S	i	s	i	i	sI	S
Mg	S	i	i	s	S	i	S	S	S	S	S	S	i	S	S	s	i	i	i	S	S
K	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
NH_4	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S		S		S	S
Na	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S

S*, soluble in water; *s*, sparingly soluble in water; *i*, insoluble in water, but made soluble by acids; *I*, insoluble in water or acids; *sI*, sparingly soluble in water, not transposed by acids; *Si*, soluble in acidulated water

ATOMIC WEIGHTS OF THE ELEMENTS.

An alphabetical list of the sixty-eight recognized elements, with their symbols and atomic weights, is here given for convenience of reference. The names of the rarer elements which are at present of little importance are printed in italics: —

Aluminium	**Al**	27.4	*Molybdenum*	**Mo**	96
Antimony	**Sb**	120	Nickel	**Ni**	58.8
Arsenic	**As**	75	*Niobium*	**Nb**	94
Barium	**Ba**	137	Nitrogen	**N**	14
Beryllium	**Be**	9	*Osmium*	**Os**	199
Bismuth	**Bi**	210	Oxygen	**O**	16
Boron	**B**	11	*Palladium*	**Pd**	106.5
Bromine	**Br**	80	Phosphorus	**P**	31
Cadmium	**Cd**	112	Platinum	**Pt**	197.4
Cæsium	**Cs**	133	Potassium	**K**	39.1
Calcium	**Ca**	40	*Rhodium*	**Rh**	104
Carbon	**C**	12	*Rubidium*	**Rb**	85.7
Cerium	**Ce**	92	*Ruthenium*	**Ru**	104
Chlorine	**Cl**	35.5	*Samarium*	**Sm**	150
Chromium	**Cr**	52.5	*Scandium*	**Sc**	44
Cobalt	**Co**	58.8	*Selenium*	**Se**	79.5
Copper	**Cu**	63.4	Silicon	**Si**	28
Didymium	**D**	95	Silver	**Ag**	108
Erbium	**E**	112.6	Sodium	**Na**	23
Fluorine	**F**	19	*Strontium*	**Sr**	87.5
Gallium	**Ga**	69	Sulphur	**S**	32
Glucinum	**Gl**	14	*Tantalum*	**Ta**	182
Gold	**Au**	196	*Tellurium*	**Te**	128
Hydrogen	**H**	1	*Terbium*	**Tb**	148.5
Indium	**In**	113.4	*Thallium*	**Tl**	204
Iodine	**I**	127	*Thorium*	**Th**	231.4
Iridium	**Ir**	198	Tin	**Sn**	118
Iron	**Fe**	56	*Titanium*	**Ti**	50
Lanthanum	**La**	93	*Tungsten*	**W**	184
Lead	**Pb**	207	*Uranium*	**Ur**	120
Lithium	**Li**	7	*Vanadium*	**V**	51.3
Magnesium	**Mg**	24	*Yttrium*	**Yt**	61.6
Manganese	**Mn**	55	Zinc	**Zn**	65
Mercury	**Hg**	200	*Zirconium*	**Zr**	89.6

APPENDIX.

To determine the empirical formula of a substance from its percentage composition.

RULE.

Divide the percentage amount of each constituent by its corresponding atomic weight; then divide each quotient so found by the lowest number, and reduce them to their simplest ratios.

EXAMPLES.

1. A body on analysis yielded the following percentage composition:—

Carbon,	27.273
Oxygen,	72.727
	100.000

Calculate its formula.

The atomic weight of carbon is 12.
The atomic weight of oxygen is 16.

Then,
$$C = \frac{27.273}{12} = 2.2727;$$
$$O = \frac{72.727}{16} = 4.5454.$$

Simplest ratio between the carbon and oxygen is as $1:2$; for
$$2.2727 : 4.5454 :: 1 : 2.$$

Hence the formula is CO_2.

2. A compound was found to have the following percentage composition:—

Nitrogen,	82.353
Hydrogen,	17.647
	100.000

Calculate its formula.

The atomic weight of nitrogen = 14, and of hydrogen = 1.

$$N = \frac{82.353}{14} = 5.882;$$

$$H = \frac{17.647}{1} = 17.647.$$

The simplest ratio between nitrogen and hydrogen is as 1 : 3;

for $5.882 : 17.647 :: 1 : 3.$

The formula of the body therefore is NH_3.

3. A compound of iron and oxygen has the following percentage composition:—

Iron,	70.01
Oxygen,	29.99
	100.00

Calculate its formula.

Atomic weight of iron 56.0, and of oxygen 16.0.

4. Deduce the formulæ of the following substances:—

Nitrogen,	30.43
Oxygen,	69.57
	100.00

5.	Potassium,	28.73
	Hydrogen,	0.73
	Sulphur,	23.52
	Oxygen,	47.02
		100.00
6.	Carbon,	20.00
	Oxygen,	26.67
	Sulphur,	53.33
		100.00

Relative Density of Solids, Liquids, and Gases; Vapor Density.

The specific gravity (sp. gr.), or relative density of a solid or liquid substance, is the ratio of its mass to the mass of an equal volume of some liquid taken as unity.

The standard universally adopted is pure water at its maximum density. The number expressing the relative density of a solid or liquid expresses how much heavier or lighter the substance is than an equal volume of water at 4° C.

The relative density of a solid is generally ascertained by the following formula, the body being first weighed in air and then in water at 4° C., and the weights carefully taken.

$$\text{Rel. dens.} = \frac{\text{Weight of substance } (W)}{\text{Weight of equal vol. of water at } 4^\circ \text{ C.}}$$

$$= \frac{W}{W - W'},$$

where $W' =$ Weight of substance in water at 4° C.

If the solid be lighter, volume for volume, than water, a sinker is attached, whose weight in water $= x$, and rel. dens. $= d$.

The relative density of the substance lighter than water is then expressed by the formula

$$\text{Rel. dens.} = \frac{W}{W + x - W''},$$

where $W'' =$ weight of sinker and solid in water.

If the relative density be required at $t°$ C., then

$$\text{Rel. dens.} = \frac{W}{W - W'} \times \text{rel. dens. of water at } t° \text{ C.}$$

The relative density of a liquid is commonly found by

1st. The specific gravity flask method.

Let $x =$ weight of flask empty, $W =$ weight of flask filled with water at $t°$ C., $W' =$ weight of flask filled with liquid under examination; then

$$\text{Rel. dens.} = \frac{W' - x}{W - x} \times \text{rel. dens. of water at } t° \text{ C.}$$

2d. By weighing a solid of constant volume in water and then in the liquid.

Let $x =$ weight of solid in air, $W =$ weight in water at $t°$ C., $W' =$ weight in liquid; then

$$\text{Rel. dens.} = \frac{x - W'}{x - W} \times \text{rel. dens. of water at } t° \text{ C.}$$

Density of Gases and Vapors.

The specific gravity, or relative density of a gas or vapor, as has been shown on pp. 18 and 19, is the ratio of its mass to the mass of an equal volume of hydrogen, measured at the same temperature and pressure. By the density of gas or vapor, we mean the relative density.

One liter of hydrogen gas at 0° C., and 760 mm. barometric pressure at the sea-level weighs .0896 gram.

The relative density of a gas is determined by weighing a known volume of the gas, and comparing it with the weight of an equal volume of hydrogen under like conditions of temperature and pressure (p. 21). The effusion method may also be used. "The rel. dens. varies directly as the square of the time of effusion of equal volumes." Graham's Law (see p. 25) may also be applied.

EXAMPLES.

1. Calculate the relative density of a solid from the following data: —

Weight of substance in air, 2.4554 grams.

Weight of substance in water, 2.0778 grams.

2. Determine the relative density of wood from the following data: —

Weight of wood in air, 4 grams.

Weight of silver sinker in air, 10 grams.

Weight of wood and sinker under water, 8.5 grams.

Relative density of silver = 10.5.

3. A solid weighs in vacuo 100 grams, in water 85 grams, and in another liquid 88 grams. What is the relative density of this liquid?

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